

## Final Revision

### Chapter (1)

#### Question (1)

- ① The Red spectral line of Li occurs at 671 nm ( $\lambda = 671 \text{ nm}$ ) calculate the energy of one photon of this light?

(Solution)

$$E = ??? \quad \lambda = 671 \text{ nm} \times 10^{-9} = 6.7 \times 10^{-7} \text{ m}$$

$$\therefore E = h \cdot \nu = \frac{h \cdot c}{\lambda} \quad \text{Since } \nu = \frac{c}{\lambda}$$

$$\nu = \frac{3.0 \times 10^8 \text{ m/sec}}{6.71 \times 10^{-7} \text{ m}} = 4.47 \times 10^{14} \frac{1}{\text{sec}}$$

$$\therefore E = h \cdot \nu = 6.63 \times 10^{-34} \text{ J} \cdot \text{sec} \times 4.47 \times 10^{14} \frac{1}{\text{sec}} = 2.96 \times 10^{-19} \text{ Joule}$$

#### Question (2)

- ② The Following are representative wavelengths in the infrared, ultraviolet and X-Ray regions of the electromagnetic spectrum respectively;  $1 \times 10^{-6} \text{ m}$ ,  $1 \times 10^{-8} \text{ m}$ ,  $1 \times 10^{-10} \text{ m}$ , what is the energy of a photon of each Radiation?

(Solution)

$$\lambda \text{ in infrared} = 1 \times 10^{-6} \text{ m} \quad \lambda \text{ in ultraviolet} = 1 \times 10^{-8} \text{ m} \quad \lambda \text{ in X-Ray} = 1 \times 10^{-10} \text{ m}$$

$$\therefore E = h \cdot \nu = \frac{h \cdot c}{\lambda} \quad \text{Since } \nu = \frac{c}{\lambda}$$

$$\text{① For infrared } \nu = \frac{3.0 \times 10^8 \text{ m/sec}}{1 \times 10^{-6} \text{ m}} = 3.0 \times 10^{14} \frac{1}{\text{sec}}$$

$$E = h \cdot \nu = 6.63 \times 10^{-34} \text{ J} \cdot \text{sec} \times 3.8 \times 10^{14} \frac{1}{\text{sec}} = 2.589 \times 10^{-19} \text{ Joules}$$

$$\text{② For ultraviolet } \nu = \frac{3.0 \times 10^8}{1 \times 10^{-8}} = 3.0 \times 10^{16} \frac{1}{\text{sec}}$$

$$E = h \cdot \nu = 6.63 \times 10^{-34} \text{ J} \cdot \text{sec} \times 3.8 \times 10^{16} \frac{1}{\text{sec}} = 2.589 \times 10^{-17} \text{ Joules}$$

$$\text{For X-Ray} = ?? \quad 2.589 \times 10^{-15} \sim \text{Complete}$$

### Question (3)

- 3) Calculate the wavelength of light emitted when the electron in a atom undergoes transition from energy level  $n=4$  to  $n=2$ ?

Solution

$$\lambda = ??? \quad n_1 = 2 \quad n_2 = 4 \quad R = 1097 \times 10^4$$

$$\therefore \frac{1}{\lambda} = 109678 \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1097 \times 10^4 \left( \frac{1}{2^2} - \frac{1}{4^2} \right) = 1097 \times 10^4 \left( \frac{1}{4} - \frac{1}{16} \right)$$

$$\therefore \frac{1}{\lambda} = 1097 \times 10^4 \times \frac{(4-1)}{16} = 1097 \times 10^4 \times \frac{3}{16} = 205.6875 \times 10^4$$

$$\therefore \frac{1}{\lambda} = 205.6875 \times 10^4 \quad \therefore \lambda = \frac{1}{205.6875 \times 10^4} = 0.0048617 \times 10^{-4} \text{ m} = 4.86 \times 10^{-7} \text{ m}$$

$$\therefore \lambda = 4.86 \times 10^{-7} \text{ m}$$

### Question (4)

- 4) Calculate the wavelength of the light emitted from hydrogen atom when the electron undergoes transition from level  $n=3$  to  $n=1$ ?

Solution

$$\lambda = ??? \quad n_1 = 1 \quad n_2 = 3 \quad R = 1097 \times 10^4$$

$$\therefore \frac{1}{\lambda} = 1097 \times 10^4 \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1097 \times 10^4 \left( \frac{1}{1^2} - \frac{1}{3^2} \right) = 1097 \times 10^4 \left( 1 - \frac{1}{9} \right)$$

$$\frac{1}{\lambda} = 1097 \times 10^4 \times \frac{8}{9} = \frac{8776}{9} \times 10^4 = 975.11111 \times 10^4$$

$$\therefore \lambda = \frac{1}{975.11111 \times 10^4} = 0.0010255 \times 10^{-4} \text{ m} = 1.02552 \times 10^{-7} \text{ m}$$

### Question (5)

- 5) Calculate the frequency of the electromagnetic radiation emitted by the hydrogen atom in the electron transition  $n=4$  to  $n=3$ ?

$$C = 3 \times 10^8 \text{ m/sec} \quad R = 1097 \times 10^4 \quad (\text{Exam June 2010})$$



### Solution

$$C = 3 \times 10^8 \text{ m/sec} \quad n_1 = 3 \quad n_2 = 4 \quad R = 1097 \times 10^4 \text{ m}^{-1} \quad \lambda = ???$$

$$\therefore \lambda = \frac{C}{\nu} \quad \text{and} \quad C = 3 \times 10^8 \text{ m/sec} \quad \text{So } \lambda \text{ is Required}$$

$$\frac{1}{\lambda} = 1097 \times 10^4 \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1097 \times 10^4 \left( \frac{1}{3^2} - \frac{1}{4^2} \right) = 1097 \times 10^4 \left( \frac{1}{9} - \frac{1}{16} \right)$$

$$\frac{1}{\lambda} = 1097 \times 10^4 \left( \frac{16-9}{144} \right) = 1097 \times 10^4 \times \frac{7}{144} = 53.326389 \times 10^4$$

$$\therefore \lambda = \frac{1}{53.326389 \times 10^4} = 0.0187524 \times 10^{-4} \text{ m}$$

$$\nu = \frac{C}{\lambda} = \frac{3 \times 10^8 \text{ m/sec}}{0.0187524 \times 10^{-4}} = 159.97952 \times 10^{12} \text{ Sec}^{-1}$$

### Question (6)

⑥. An electron in a hydrogen atom in the level  $n=5$  undergoes a transition to level  $n=3$  what is the frequency of the emitted  
 $R = 1097 \times 10^4 \text{ m}^{-1}$ ,  $C = 3 \times 10^8 \text{ m/sec}$

### Solution

$$C = 3 \times 10^8 \text{ m/sec} \quad n_1 = 3 \quad n_2 = 5 \quad R = 1097 \times 10^4 \text{ m}^{-1} \quad \lambda = ???$$

$$\therefore \lambda = \frac{C}{\nu} \quad \text{and} \quad C = 3 \times 10^8 \text{ m/sec} \quad \text{So } \lambda \text{ is Required}$$

$$\frac{1}{\lambda} = 1097 \times 10^4 \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1097 \times 10^4 \left( \frac{1}{3^2} - \frac{1}{5^2} \right) = 1097 \times 10^4 \left( \frac{1}{9} - \frac{1}{25} \right)$$

$$\therefore \frac{1}{\lambda} = 1097 \times 10^4 \left( \frac{25-9}{225} \right) = 1097 \times 10^4 \times \frac{16}{225} = 780088.88$$

$$\therefore \lambda = \frac{1}{780088.88} = 1.281905 \times 10^{-6} \text{ m}$$

$$\nu = \frac{C}{\lambda} = \frac{3 \times 10^8}{1.281905 \times 10^{-6}} = 2.3402 \times 10^{14} \text{ Sec}^{-1}$$

### Question (7)

⑦. Calculate the wavelength of the first line of Lyman Series of the hydrogen atom spectrum?  $R = 1097 \times 10^4 \text{ m}^{-1}$

## Solution

In the Lyman Series  $n_1 = 1$   $n_2 = 2$  (First line)

$$\frac{1}{\lambda} = 1097 \times 10^4 \left( \frac{1}{1^2} - \frac{1}{2^2} \right) = 1097 \times 10^4 \left( \frac{4-1}{4} \right)$$

$$\frac{1}{\lambda} = 1097 \times 10^4 \times \frac{3}{4} = 8227500$$

$$\therefore \lambda = \frac{1}{8227500} = 1.21543 \times 10^{-7} \text{ m}$$

## Question (8)

A line in the Paschen Series of the emission Spectrum of Hydrogen is observed at wave number =  $7800 \text{ cm}^{-1}$ , Deduce the upper State Principal quantum number for this transition where  $R = 109729 \text{ cm}^{-1}$  (Exam Jan 2010)?

## Solution

$$\text{Wave number } \bar{\nu} = \frac{1}{\lambda} = 7800 \text{ cm}^{-1} \quad [n_1 = 3 \text{ (Paschen Series)}]$$

$$S_{n_2} = ??? \quad R = 109729 \text{ cm}^{-1}$$

$$\frac{1}{\lambda} = 109729 \text{ cm}^{-1} \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

$$7800 \text{ cm}^{-1} = 109729 \left( \frac{1}{3^2} - \frac{1}{n_2^2} \right)$$

$$\frac{7800}{109729} = \frac{1}{9} - \frac{1}{n_2^2} \quad (\text{in Paschen } n_1 = 3)$$

$$0.07108 - \frac{1}{9} = -\frac{1}{n_2^2} \quad \therefore -0.040031 = -\frac{1}{n_2^2}$$

$$\therefore n_2^2 = \frac{1}{0.040031} = 24.98 \approx 25$$

$$n_2 = \sqrt{25} = 5 \quad \therefore n_2 = 5$$



### Question 19)

Q. write a scientific term for the following :-

1. The number of peaks that pass a given point in one second as the wave moves (Frequency  $\nu$ ) (exam Jan 2013)

2. Maximum height of wave peak (amplitude)

3. Distance between consecutive peaks or troughs in the wave (wavelength  $\lambda$ ) (exam Jan 2013)

### Question 110)

Q. Choose the correct answer: (Exam Jan 2010)

\* He can be singly ionized by losing one electron to become  $\text{He}^+$  cation which of the following statements is true concerning this helium cation?

a) The line spectrum of this helium cation will resemble the line spectrum of Lithium cation

b) The line spectrum of Helium cation will remain the same as for unionized helium

c) The line spectrum of  $\text{He}^+$  cation will resemble the line spectrum of Hydrogen atom (✓)

d) The line spectrum of  $\text{He}^+$  cation will resemble the line spectrum of Li

→ Solution: True answer is (C)

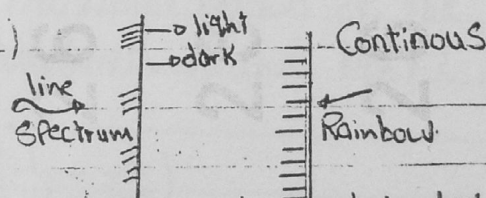
### Question 11)

11) Define the following:-

1a) Continuous and line spectrum

\* Continuous spectrum:- Rainbow that you see when white light passes through a prism, it includes all of visible light wavelengths from 380 nm to 780 nm.

\* Line spectrum:- Shows only colours that are emitted from a substance. So rather than seeing all of the wavelengths, you would see only those associated with material. Each compound has its own spectra (Bright and dark areas are seen)



$$\lambda = \frac{h}{p}$$

### Question (12)

\* Determine the DeBroglie wavelength of

①. An electron moving at 10 km/h or one tenth the speed of light

②. A 400 gm Frisbee moving at 10 km/h.

### Solution

De Broglie equation states that  $\lambda = \frac{h}{p} = \frac{h}{m \cdot v}$

where  $h = 6.63 \times 10^{-34} \text{ J} \cdot \text{sec}$

$m = 9.11 \times 10^{-31} \text{ kg}$

$$v = \frac{1}{10} \times 3 \times 10^8 = 3 \times 10^7 \text{ m/sec}$$

$$\therefore \lambda = \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{sec}}{9.11 \times 10^{-31} \text{ kg} \times 3 \times 10^7 \text{ m/sec}} = 2.4 \times 10^{-11} \text{ m}$$

②. Frisbee (400 gm) moving at 10 km/h

$$v = \frac{10 \text{ km}}{3600} = \frac{10 \times 1000}{3600} = 2.77 \text{ m/sec}$$

$$\lambda = \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{sec}}{0.400 \text{ kg} \times 2.77 \text{ m/sec}} = \frac{6.63 \times 10^{-34}}{1.108} = 6.63 \times 10^{-34} \text{ m}$$

### Question (13)

\* The transition from the  $n=7$  to the  $n=2$  level of the hydrogen atom is accompanied by the emission of light slightly beyond the range of human perception in the ultraviolet region of the spectrum. Determine the energy and wavelength of this light.

### Solution

$$\frac{1}{\lambda} = 1097 \times 10^4 \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1097 \times 10^4 \left( \frac{1}{2^2} - \frac{1}{7^2} \right)$$

$$= 1097 \times 10^4 \left( \frac{1}{4} - \frac{1}{49} \right) = 1097 \times 10^4 \left( \frac{49-4}{196} \right) = 1097 \times 10^4 \times 0.2295$$

$$\frac{1}{\lambda} = 2518622.449$$

$$\therefore \lambda = \frac{1}{2518622.449} = 3.9704 \times 10^{-7} \text{ m}$$

$$E = \frac{h \cdot c}{\lambda} = \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{sec} \times 3 \times 10^8 \text{ m/sec}}{3.97 \times 10^{-7} \text{ m}} = 5.009 \times 10^{-19} \text{ Joules}$$



### Question (14) ✓

- (14) using the equation  $E = R_H \left( \frac{1}{2^2} + \frac{1}{n^2} \right)$  determine the energies and wavelengths of visible emission bands in the atomic spectrum of Hydrogen arising from  $n=4, 5$ , and 6

Solution

For  $n=4$

$$\therefore E = R_H \left( \frac{1}{2^2} + \frac{1}{4^2} \right) =$$

$$E = 2.18 \times 10^{-18} \text{ Joules} \left( \frac{1}{4} + \frac{1}{16} \right) = 2.18 \times 10^{-18} \left( \frac{4+1}{16} \right) = 2.18 \times 10^{-18} \times \frac{5}{16}$$

$$= 6.8125 \times 10^{-19} \text{ Joules}$$

$$E = h \cdot \frac{c}{\lambda} \quad \therefore \lambda = \frac{h \cdot c}{E} = \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{sec} \times 3 \times 10^8 \text{ m/sec}}{6.8125 \times 10^{-19} \text{ Joules}}$$

$$= 2.91963 \times 10^{-7} \text{ m (complete for } n=5 \text{ to } 6)$$

### Question (15) ✓

- (15) Calculate the atomic weight of (Cr) from isotope mass & fractional abundance?

Isotopes for the same element: Nuclei having the same number of protons (atomic number) but differ in number of neutrons (mass number)

For Cr: Four isotopes are available

Isotope	Mass	abundance
$\text{Cr}^{50}$	49.946 amu	4.345 %
$\text{Cr}^{52}$	51.940 amu	83.789 %
$\text{Cr}^{53}$	52.940 amu	9.501 %
$\text{Cr}^{54}$	53.938	2.365 %

Atomic weight = average mass of the atoms of element  
 $\times \frac{\text{Relative abundance}}{100}$

$$= \left( \frac{49.946 \times 4.345}{100} \right) + \left( \frac{51.940 \times 83.7}{100} \right) + \left( \frac{52.940 \times 9.501}{100} \right) + \left( \frac{53.938 \times 2.365}{100} \right)$$
$$= 51.996 \text{ amu}$$

Very important  
wavelength ( $\lambda$ ) units

- ① metre
- ② Centimeter =  $0.01 \text{ m} = 10^{-2} \text{ m}$
- ③ millimeter =  $0.001 \text{ m} = 10^{-3} \text{ m}$
- ④ micrometer =  $10^{-6} \text{ m}$
- ⑤ nanometer =  $10^{-9} \text{ m}$
- ⑥ Angstrom =  $10^{-10} \text{ m}$
- ⑦ Picometer =  $10^{-12} \text{ m}$

with my Best wishes.

E. Shoukry



## Final Revision

### CHAPTER (2)

#### Question (1)

①. What is the Complete Symbols for these atoms:-

a) it Contains 17 Protons and 18 Neutrons in it's Nucleus?

b) it Contains 11 electrons and 12 Neutrons in it's Nucleus?

c) it Contains 38 Protons and 50 Neutrons in it's Nucleus?

#### Solution

a) atom is Chloride ( $\text{Cl}$ ) with electronic Configuration:  $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

b) atom is Sodium ( $\text{Na}$ ) with electronic Configuration:  $1s^2, 2s^2, 2p^6, 3s^1$

c) atom is Strontium ( $\text{Sr}$ ) with electronic Configuration:  $\text{Kr } 5s^2$

#### Question (2)

②. Give the number of Protons and neutrons in each of the following :-  $\text{Be}^4_2$ ,  $\text{Fe}^{56}_{26}$ ,  $\text{Zn}^{64}_{30}$

#### Answer

a)  $\text{Be}^4_2$  : it Contains (2) Protons and (2) neutrons.

2)  $\text{Fe}^{56}_{26}$  : it Contains (26) Protons and (30) Neutrons.

3)  $\text{Zn}^{64}_{30}$  : it Contains (30) Protons and (34) Neutrons.

#### Question (3)

③. which of the following orbital diagram or electron Possible and which are impossible according to the Pauli exclusion

#### Answer

a)  $[1\uparrow][1\downarrow][\uparrow]$  (Possible)  $\rightarrow$  there is No more than  $2e^-$  in any orbital and Spinning Motion of electrons is Correct.

b)  $[1\uparrow][\uparrow\downarrow][1\downarrow]$  (impossible)  $\rightarrow$  Three electrons are in a  $2s$  orbital and two of them are with the Same Spin

c)  $[1\uparrow][1\uparrow][1\uparrow\uparrow]$  (impossible) because two electrons with the same spin in a 2p orbital.

d)  $1s^3, 2s^1$  (impossible)  $\rightarrow$  Presence of three electrons in 1s subshell (Maximum occupation for 1s =  $2e^-$ )

e)  $1s^2, 2s^1, 2p^7$  (impossible)  $\rightarrow$  Presence of seven electrons in 2p subshell (Maximum occupation for 2p =  $6e^-$ )

f)  $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^8, 4s^2$  (Possible)  $\rightarrow$  Maximum occupation of 3d orbital =  $10e^-$  so possible electronic configuration.

### Question (4)

4) use the building up principle to obtain the electron configuration for the ground state of: 1) gallium (Z=31) 2) Mn (Z=25) 3) Al (Z=13) 4) gadolinium (Z=64)  $\rightarrow$  Gd (No abbreviate for the core)

### Answer

The building up order of subshells according to increase in energy is:

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6,$

$6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^2, 5f^{14}, 6d^{10}, 7p^6, 6f^{14}, 7d^{10}, 7f^{14}$

So according to this order the electron configuration of these elements will be:

1) Ga  $\rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^1$  (Period (4)) (Group 3A)

2) Mn  $\rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^5$  (Period (4)) (Group 7B)

3) Al  $\rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$  (Period (3)) (Group 6A)

4) Al  $\rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$  (Period (3)) (Group 3A)

5) Gd  $\rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^7, 5d^1$   
or  $[Xe] 4f^7 5d^1 6s^2$  (lanthanide element)



### Question (5)

⑤ \* Write the orbital diagram for the ground state of:-

1. Cobalt ( $Z=27$ )

(2) - Magnesium ( $Z=12$ )

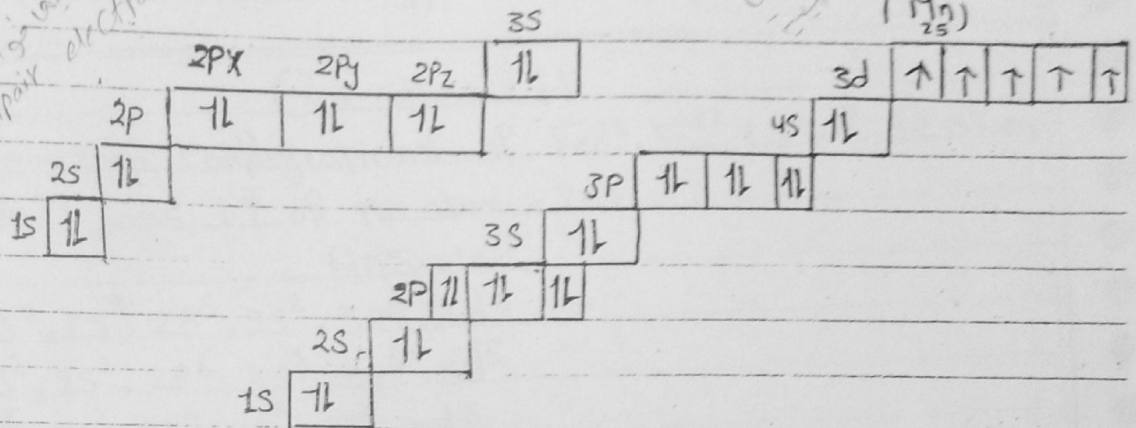
(3) - Cerium ( $Z=58$ )

is the atom diamagnetic or Paramagnetic?

\* Answer \*

① - Mg

*Significant  
No unpaired electron*



(3) - Cerium  $\text{Ce} \rightarrow (\text{Xe}) 4f^1, 5d^1, 6s^2$  (Fill by yourself in diagram)

(B) \* Is the atom Paramagnetic or diamagnetic?

→ An atom is diamagnetic if all of its electrons are paired

→ An atom is Paramagnetic if it has one or more unpaired electron

So that:- 1. Mg is diamagnetic (All electrons are paired)

2. Mn, Ce are Paramagnetic because there is unpaired electrons in Both two elements.

### Question (6)

⑥ \* Which of the following electron Configurations are Possible ones? Explain why the others aren't?

Answer

a)  $1s^2 2s^2 2p^7$  (impossible) because 2p has 7e and it can hold only 6e only

b)  $1s^2, 2s^2, 3d^{11}$  (impossible) because 3d orbital has 11e and it can hold only 10e (3d isn't after 2s according to energy diagram)

c)  $1s^2, 2s^2, 2p^5 \rightarrow$  (Possible)

d)  $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{12} \rightarrow$  (impossible) 3d orbital can't have  $12e^-$  (can hold  $10e^-$  only)

e)  $1s^2, 2s^2, 2p^6, 3s^2, 3p^7, 3d^9 \rightarrow$  (impossible) 3p orbital can't have  $7e^-$  (can hold  $6e^-$  only)

### Question (7)

Q. write the electron configurations of  $Fe^{+3}, Fe^{+2}$  (Z of Fe atom = 26) and  $Mn^{+2}$  ions (Z of Mn atom = 25)

Answer

a)  $Fe^{+2} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^6$

b)  $Fe^{+3} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5$

c)  $Mn^{+2} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5$

Note: electrons of 4s are lost first followed by 3d electrons

### Question (8)

Q. Give the electronic configurations of  $Ni^{+2}, Ni^{+3}, Cu^+, Cu^{+2}, Sn, Sn^{+2}, Bi, Bi^{+3}$  where the atomic numbers are: a)  $Ni \rightarrow Z=28$   
 $Cu \rightarrow Z=29$  &  $Bi \rightarrow Z=83$  (Make by yourself for Br, V,  $Hg^{+2}$ )

Answer

1)  $Ni^{+2} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^8$  (two electrons are lost from 4s)

2)  $Ni^{+3} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^5$  (Keeping 3d half filled and 4s completely filled for reaching the most stable state)

3)  $Cu^+ \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10} \rightarrow [Ar]_{18} 3d^{10}$

4)  $Cu^{+2} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10} \rightarrow [Ar]_{18} 3d^9$

5)  $Sn \rightarrow [Kr] 5s^2, 4d^{10}, 5p^2$

6)  $Sn^{+2} \rightarrow [Kr]_{36} 5s^2, 4d^{10}$

7)  $Bi \rightarrow [Xe]_{54} 6s^2, 4f^{14}, 5d^{10}, 6p^3$

8)  $Bi^{+3} \rightarrow [Xe]_{54} 6s^2, 4f^{14}, 5d^{10}$



## \* Atomic Quantum Numbers \*

\* How to calculate Four or Possible quantum numbers?

\* For Calculating all Possible values for different quantum numbers?

$n$	$L (0 \rightarrow n-1)$	No of values = $(2L+1)$ $ML$ القيم الممكنة لـ $L$	$m_s$
1	0	0	$+\frac{1}{2}$ or $-\frac{1}{2}$
2	0, 1	-1, 0, 1	$+\frac{1}{2}$ or $-\frac{1}{2}$
3	0, 1, 2	-2, -1, 0, 1, 2	$+\frac{1}{2}$ or $-\frac{1}{2}$
4	0, 1, 2, 3	-3, -2, -1, 0, 1, 2, 3	$+\frac{1}{2}$ or $-\frac{1}{2}$

\* Note that:- In determination of Subshells ( $L$ ) you must know that

$L$	Subshell
0	S
1	P
2	d
3	f

\* Applying this to Calculate Four Possible quantum numbers for any electron we can use the following:-

$n$	$L$	Subshell	$ML$	$m_s$
level (1)	0 $\rightarrow$ 1S	1S	$\boxed{\uparrow\downarrow}$ (0)	one electron = $+\frac{1}{2}$ & other = $-\frac{1}{2}$
level (2)	0, 1	0 $\rightarrow$ 2S 1 $\rightarrow$ 2P	2S: $\boxed{\uparrow\downarrow}$ (0) 2P: $\boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow}$ -1 0 1	$+\frac{1}{2}$ & $-\frac{1}{2}$
level (3)	0, 1, 2	0 $\rightarrow$ 3S 1 $\rightarrow$ 3P 2 $\rightarrow$ 3d	3S: $\boxed{\uparrow\downarrow}$ (0) 3P: $\boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow}$ -1 0 1 3d: $\boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow} \boxed{\uparrow\downarrow}$ -2 -1 0 1 2	$+\frac{1}{2}$ & $-\frac{1}{2}$

n	l	Subshell	ml
4s	0, 1, 2, 3	0 → 4s 1 → 4p 2 → 4d 3 → 4f	4s: $\uparrow\downarrow$ (0) 4p: $\uparrow\downarrow$ (-1), $\uparrow\downarrow$ (0), $\uparrow\downarrow$ (1) 4d: $\uparrow\downarrow$ (-2), $\uparrow\downarrow$ (-1), $\uparrow\downarrow$ (0), $\uparrow\downarrow$ (1), $\uparrow\downarrow$ (2) 4f: $\uparrow\downarrow$ (-3), $\uparrow\downarrow$ (-2), $\uparrow\downarrow$ (-1), $\uparrow\downarrow$ (0), $\uparrow\downarrow$ (1), $\uparrow\downarrow$ (2), $\uparrow\downarrow$ (3)

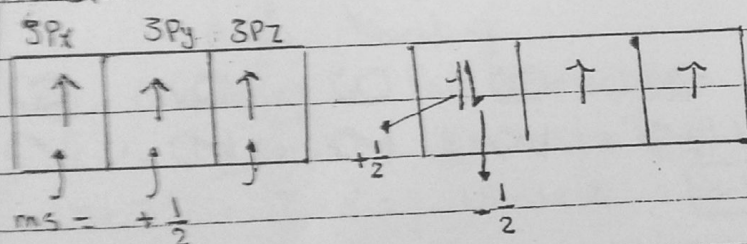
So it is easy to determine  $n, l, m_l, m_s$  for any electron inside the atom.

For  $m_s$  it will be either  $+\frac{1}{2}$  or  $-\frac{1}{2}$ .

when electron direction inside orbital is spin up  $\uparrow m_s = +\frac{1}{2}$

and when spin down  $\downarrow m_s = -\frac{1}{2}$

We should know that electrons are filled in a single way in orbitals in spin up direction then we begin to pair as follows:-



### Question (9)

Q. which one of the following sets of Quantum numbers  $(n, l, m_l, m_s)$  represents as impossible arrangement?

\* Answer \*

1)  $3, 2, -2, +\frac{1}{2}$  → Possible because  $n=3, l=2$  (3d),  $m_l=-2$  (Possible Value),  $m_s = +\frac{1}{2}$  (Possible)

2)  $4, 0, 0, \frac{1}{2}$  → Possible as  $n=4, l=0$  (4s),  $m_l=0$  (Possible)  
 $m_s = +\frac{1}{2}$  (Possible)



## \* Slater Rule \*

### Question (20)

(20) \* Using Slater's Rule, Calculate  $Z_{\text{eff}}^*$  for

1. 3p electron in P, S, Cl and Ar? Is the calculated values of  $Z^*$  consistent with the relative sizes of these atoms?

Answer (Add Al, Si and answer in the

①  $\rightarrow \text{P}_{15} \rightarrow (\text{Ne})_{10}, 3s^2, 3p^3$  Same way)

①. electronic configuration of P is  $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

②. Putting electronic configuration into groups will be:

$(1s)^2, (2s, 2p)^8, (3s, 3p)^5$

③.  $Z^*$  for the last 3p electron

$$Z^* = Z - \sigma \quad \text{where } (Z = 15)$$

$\downarrow$                        $\downarrow$   
 At. Nu              Shielding

④. Calculating  $\sigma$  for 3p electron:

$$\sigma = (4 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 10.2$$

$$\text{So } Z^* = Z - \sigma = 15 - 10.2 = \boxed{4.8}$$

②  $\rightarrow \text{S}_{16} \rightarrow$  ①  $\rightarrow \text{S}_{16} : 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$   
 ②  $\rightarrow (1s)^2, (2s, 2p)^8, (3s, 3p)^6$

$$\sigma \text{ for last 3p e } \rightarrow 6 = (5 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 10.55$$

$$\text{④} \rightarrow Z^* = Z - \sigma = 16 - 10.55 = \boxed{5.45}$$

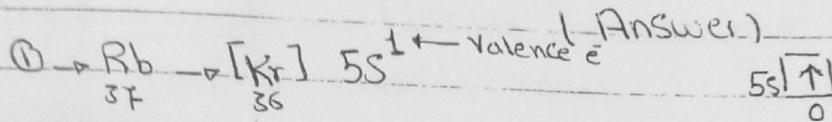
③  $\rightarrow \text{Cl}_{17} \rightarrow$  ①  $\rightarrow \text{Cl}_{17} : 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$   
 ②  $\rightarrow (1s)^2, (2s, 2p)^8, (3s, 3p)^7$

$$\sigma \text{ for last 3p e } \rightarrow 7 = (6 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 10.9$$

$$\text{④} \rightarrow Z^* = 17 - 10.9 = \boxed{6.1}$$

### Question (15)

15) \* Correct set of four quantum numbers for the valence (outmost) electron of rubidium ( $Z=37$ ) is: ?



So four quantum numbers are:  $5, 0, 0, \frac{1}{2}$

### Question (16)

16) \* Bohr model can explain: ?

\* Answer \*

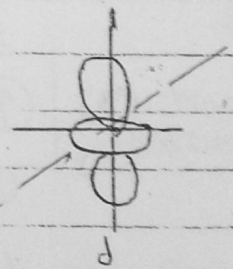
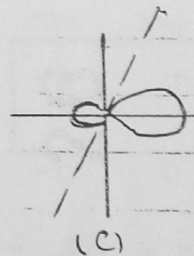
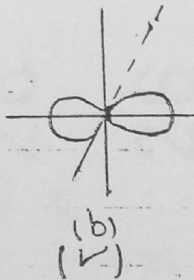
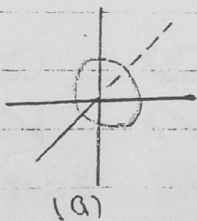
Spectrum of an atom or ion containing one electron only

### Question (17)

17) \* Choose the correct answer:-

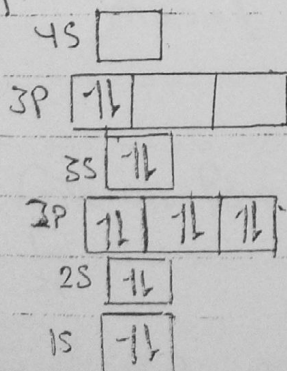
1) \* The Probability area that best represents the shape of an atomic 3p orbital is (June 2012)

a)

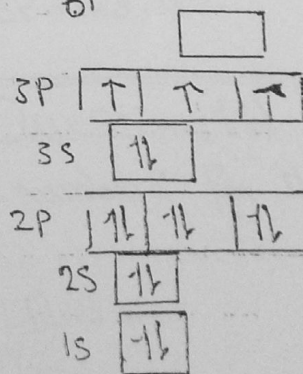


2) Which of the following orbital box diagrams represents Silicon which has 14 electrons? Si ( $Z=14$ ) (Exam June 2012)

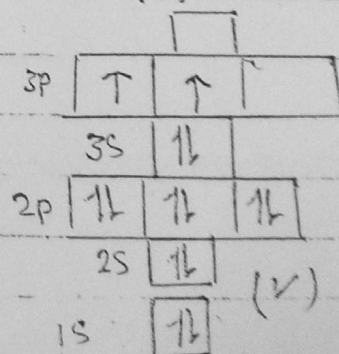
a)



b)



(c)





3)  $3, 2, -3, \frac{1}{2}$  → impossible because  $n=3$ ,  $l=2$  (3d) &  $m_l = -3$  (not possible value), so this is impossible arrangement.

4)  $5, 3, 0, -\frac{1}{2}$  → Possible because  $n=5$  &  $l=3$  (5f),  $m_l = -3$  (possible) and  $m_s = -\frac{1}{2}$  (possible value)

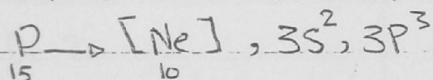
So impossible arrangement  $3, 2, -3, \frac{1}{2}$

### Question (10)

(10) Give all four quantum numbers for the valence electron in  ${}_{15}P$  atom?

\* Answer \*

① First write out electronic configuration as follows:-



② Determine the position of electron

→ It is the last electron in 3p orbital (valence electron)

3p <sub>x</sub>	3p <sub>y</sub>	3p <sub>z</sub>	
↑	↑	↑	electron position
-1	0	+1	

So  $n=3$  Shell No 3,

$l=1$  Since  $0 \rightarrow 3s$  &  $1 = 3p$  &  $2 \rightarrow 3d$  and electron

locates at 3p

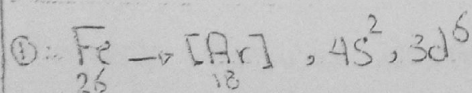
$m_l = +1$

$m_s = +\frac{1}{2}$  it is Spinned up.

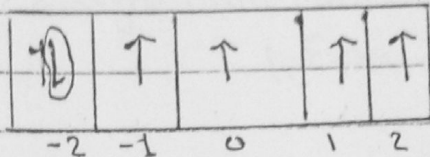
### Question (11)

(11) Give all four quantum numbers for the valence electron in  ${}_{26}Fe$ ?

Answer



2. determine electron Position



③ Four quantum numbers are:

$n=3$   $l=2$  ( $3d$ )  $m_l=-2$   $m_s=-\frac{1}{2}$  electron spin is down

### Question (12)

⑫ Find the possible values for the  $n, l, m_l, m_s$  quantum numbers for 5d electron & ( $3d, 4f$ ) ?

Answer

For 3d electron	5d electron	4f
$n=3$	$n=5$	$n=4$
$l=2$	$l=2$	$l=3$
$m_l=-2, -1, 0, 1, 2$	$m_l=-2, -1, 0, 1, 2$	$m_l=-3, -2, -1, 0, 1, 2, 3$
$m_s=+\frac{1}{2}$ or $-\frac{1}{2}$	$m_s=+\frac{1}{2}$ or $-\frac{1}{2}$	$m_s=+\frac{1}{2}$ or $-\frac{1}{2}$

### Question (13)

• which  $+2$  ion has five 3d electrons? which one has two 3d electrons?

Answer

- a)  $+2$  ion is  $Mn^{+2} \rightarrow [Ar] 3d^5$   
 b)  $+2$  ion  $2(3d)e \rightarrow Ti^{+2} \rightarrow [Ar] 3d^2$

### Question (14)

⑭ Choose the correct answer:

- a) the Principal quantum number of the atom is related to:- (Jan 2013)  
 1) Size of the orbital ( $\checkmark$ )      2) Spin angular momentum  
 3) orbital angular momentum      4) orientation of orbital in space

Answer

Correct Answer is No (1)  $\rightarrow$  Size of the orbital.



3) which of the following electron configurations represent the electron configuration for magnesium cation  $Mg^{2+}$  .

a)  $1s^2, 2s^2, 2p^6, 3s^2$

(Exam Jan 2011)

b)  $1s^2, 2s^2, 2p^6, 3s^2, 3p^2$

c)  $1s^2, 2s^2, 2p^6$  (✓)

d)  $1s^2, 2s^2, 2p^4$

4) How many valence electrons does an oxygen atom have?

a) 2

$1s^2, 2s^2, 2p^6, 3s^2, 3p^4$

(Exam Jan 2011)

b) 6 (✓)

c) 8

d) 16

Answer

(1)  $\rightarrow$  b

(2)  $\rightarrow$  c

3  $\rightarrow$  (c)

4  $\rightarrow$  (b)

Question (18)

18) Is the following set of Quantum numbers possible or not?

Explain

1)  $n=1, l=0, m_l=0, m_s=+1$

(Jan 2013)

2)  $n=3, l=2, m_l=3, m_s=-\frac{1}{2}$

3)  $n=3, l=1, m_l=-2, m_s=-\frac{1}{2}$

Answer

1)  $\rightarrow$  impossible  $m_s$  can't be  $+1$  ( $+\frac{1}{2}$  or  $-\frac{1}{2}$ )

2)  $\rightarrow$  impossible  $m_l$  can't be  $=3$  as  $l=2$  so  $m_l = -2, -1, 0, 1, 2$

3)  $\rightarrow$  impossible  $n=3, l=1$  ( $3p$ ),  $m_l = -2$  (impossible value) it can be

$-1, 0, 1$  &  $m_s = -\frac{1}{2}$  (possible)

Question (19)

Write the scientific term for the following: (Jan 2013)

1) No two electrons in an atom can have the same four

quantum numbers  $\rightarrow$  Answer  $\rightarrow$  Pauli exclusion principle

$$\textcircled{1}: 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$$

$$\textcircled{2}: (1s)^2, (2s, 2p)^8, (3s, 3p)^8$$

$$\textcircled{3}: \sigma = (7 \times 0.35) + (8 \times 0.85) + (2 \times 1) = \boxed{11.25}$$

\* Calculated values for  $Z^*$  are

	P 15	S 16	Cl 17	Ar 18
$Z_{\text{eff}}$	4.80	5.45	6.1	6.75

This shows that  $Z_{\text{eff}}$  increases along the period and attains maximum value at noble gas which account for decreasing atomic radius along the period with minimum at the noble gas.

All of P, S, Cl, Ar locate in the same period (3rd period) according to their electronic configuration

\* Across a period from left to right atomic radius decreases with increasing  $Z_{\text{eff}}$  and as from P to Ar  $Z_{\text{eff}}$  increases so values of  $Z_{\text{eff}}$  is consistent with size of atoms.

### Question (20) \*

Q. Using Slater rule Calculate  $Z^*$  for 4s, 3d electron of Cu ✓  
? Which type of electrons is more likely to be lost when Cu forms positive ion?

(Answer)

1. electronic configuration of Cu  $\rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^{10}$

2. Putting into groups:  $(1s)^2, (2s, 2p)^8, (3s, 3p)^8, (3d)^{10}, (4s)^1$

③ For 4s electron

$$Z^* = Z - \sigma$$

For 3d electron

$$Z^* = Z - \sigma$$

$$\sigma = (1 \times 0.35) + (18 \times 0.85) + (1 \times 1) = 25.3$$

$$\sigma = (9 \times 0.35) + (18 \times 1) = 21.15$$

$$Z^* = 29 - 25.3 = \boxed{3.7}$$

$$Z^* = 29 - 21.15 = \boxed{7.85}$$

We can conclude that 4s electron experiences lower  $Z^* = 3.7$  than 3d electron and is therefore likely to be lost than 3d electron when the Cu ionized.



## Question (21)

21) using Slater rule Calculate  $Z^*$  For: 2p electron in  $O^{2-}$ ,  $F^-$ ,  $Na^+$ ,  $Mg^{+2}$ . Is the Calculated values of  $Z^*$  consistent with the relative size of these atoms? (Exam June 2012)

(Answer)

①  $O^{2-} \rightarrow 1s^2, 2s^2, 2p^6$

③  $Na^+ \rightarrow 1s^2, 2s^2, 2p^6$

②  $F^- \rightarrow 1s^2, 2s^2, 2p^6$

④  $Mg^{+2} \rightarrow 1s^2, 2s^2, 2p^6$

a) All Ions ( $O^{2-}$ ,  $F^-$ ,  $Na^+$ ,  $Mg^{+2}$ ) are isoelectronic (have the same electronic Configuration of Ne  $\rightarrow 1s^2, 2s^2, 2p^6$ )

b)  $\sigma'$  Calculation For 2p electron will be the same For all ions the only difference will be in  $Z^*$  according to difference in atomic No.

c) For  $O^{2-}$ ,  $F^-$ ,  $Mg^{+2}$ ,  $Na^+$

1) electron Configuration ( $1s^2, 2s^2, 2p^6$ )

2) Putting into groups ( $1s^2$ ), ( $2s, 2p$ )<sup>8</sup>

3) Calculating  $\sigma'$ :  $\sigma'_{(2p) \text{ for anion}} = (7 \times 0.35) + (2 \times 0.85) = 4.15$

$\rightarrow$  For  $O^{2-}$

$Z^* = Z - \sigma' = 8 - 4.15 = \boxed{3.85}$	$F^-$	$Na^+$	$Mg^{+2}$
$Z^* = 9 - 4.15 = \boxed{4.85}$	$Z^* = 11 - 4.15 = \boxed{6.85}$	$Z^* = 12 - 4.15 = \boxed{7.85}$	

$$O^{2-} < F^- < Na^+ < Mg^{+2}$$

3.85      4.85      6.85      7.85       $Z_{eff}$

\* A larger  $Z_{eff}$  indicates that there is a stronger attractive force between an outer electron and the positive core, with this in mind the atomic radius should decrease with increasing  $Z_{eff}$  so that these values are consistent with size of the ions

$$Mg^{+2} < Na^+ < F^- < O^{2-} \rightarrow Z_{eff} \propto \frac{1}{\text{Radius}}$$

atomic Radius increases in this direction

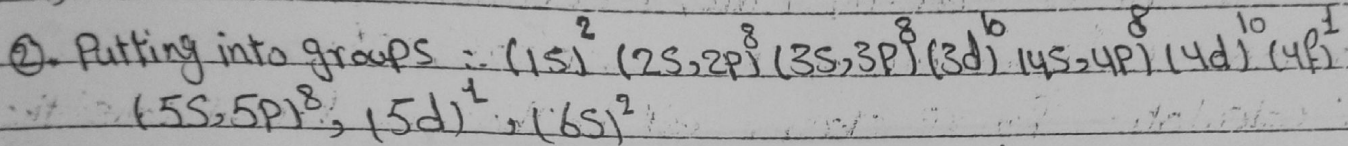
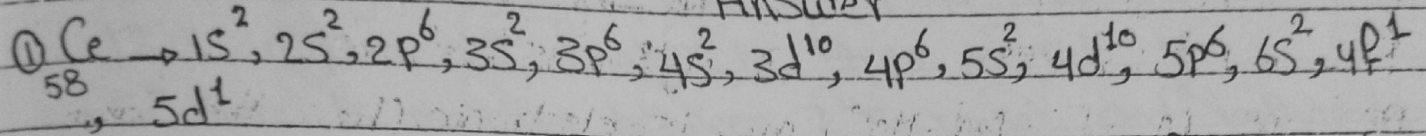
⊗  $Na^+$ ,  $Mg^{+2}$  locate in Period (3) Radius decrease across a period & Radius of positive ion < it's Neutral atom due to decreasing Repulsion force  
So  $Mg^{+2} < Na^+$

+  $O^{2-}$ ,  $F^-$  <sup>the least</sup> locate in Period (2) Radius decrease across a period, Radius of negative ion > it's Neutral atom due to increasing Repulsion force so  $O^{2-} > F^-$  and final order  $O^{2-} > F^- > Na^+ > Mg^{+2}$

## Question (22)

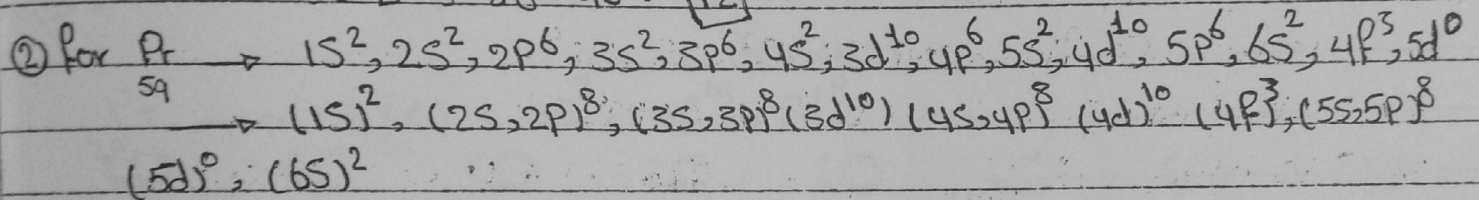
2. Calculate  $Z^*$  for: 4f electron in Ce, Pr, Nd. There is a decrease in size known as lanthanide contraction with increasing atomic number. Are your values consistent with this trend?

Answer



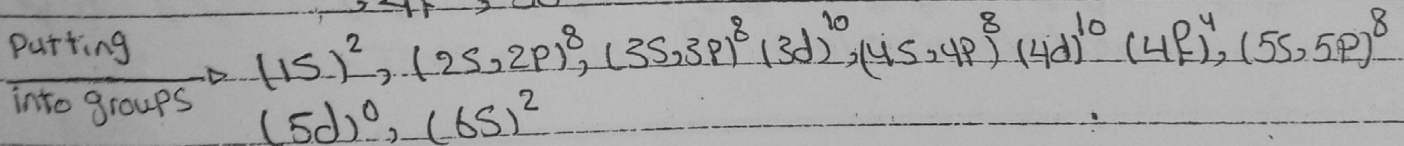
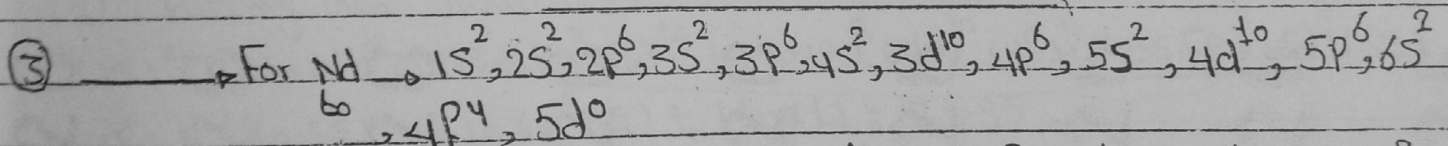
For 4f electron:  $\sigma = (0 \times 0.35) + (46 \times 1) = 46$

$Z^* = 58 - 46 = \boxed{12}$



For 4f e  $\rightarrow \sigma = (2 \times 0.35) + (46 \times 1) = 46.7$

$Z^* = 59 - 46.7 = \boxed{12.3}$



$\sigma = (3 \times 0.35) + (46 \times 1) = 47.05$

$Z^* = 60 - 47.05 = \boxed{12.95}$

Ce	Pr	Nd
58	59	60
12	12.3	12.95

The outermost electrons on going from Ce to Nd experience an increasing in  $Z^*$  and therefore are drawn into slightly closer distance from nucleus with increasing  $Z$  and  $Z^*$  and decreasing radius.



Lanthanide Contraction: decrease or smaller decrease other than expected in ionic radii of elements of lanthanide due to very poor shielding by 4f electrons so final effect for this contraction is slight decrease in radius of elements.

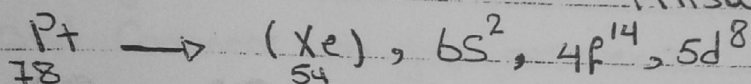
→ As Radius decreases slightly so increase in  $Z_{eff}$  will be also slightly. (This is true as it goes from 12 in Ce to 12.3 in Pr and finally to 12.95 in Nd)

→ So there is matching between  $Z^*$  increase and lanthanide contraction.

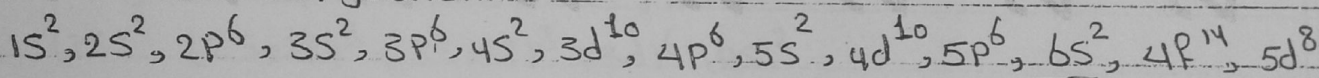
### Question (23)

(23) + Calculate  $Z^*$  for 6s electron in Platinum ( $Z = 78$ )

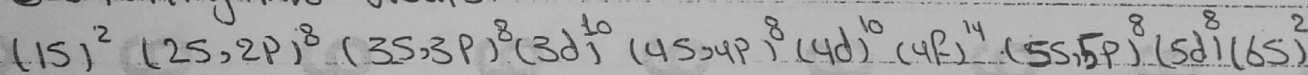
Answer



①. electronic Configuration



②. Putting into groups:-



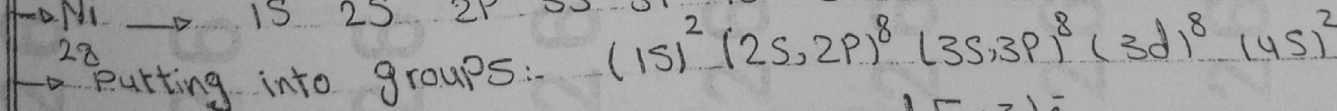
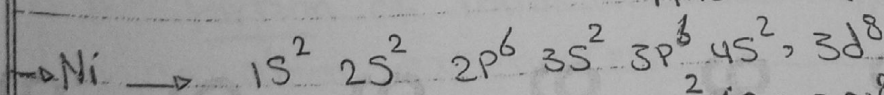
$$\sigma = (1 \times 0.35) + (16 \times 0.85) + (60 \times 1) = 73.45$$

$$Z^* = 78 - 73.45 = 4.15 \text{ For valence electron}$$

### Question (24)

(24) + Using Slater rule Calculate  $Z^*$  for 3d, 4s in Ni  $\rightarrow Z = 28$  and Zn  $\rightarrow Z = 30$

Answer (24)



For 4s e

$$\sigma = (1 \times 0.35) + (16 \times 0.85) + (10 \times 1) = 23.45$$

$$Z_{eff}^* = Z - \sigma = 28 - 23.45 = 4.05$$

For 3d e

$$\sigma = (7 \times 0.35) + (18 \times 1) = 20.45$$

$$Z^* = Z - \sigma = 28 - 20.45 = 7.55$$

4s electron experience low  $Z^*$  so it will be lost first when Ni ionizes.

$$\text{Zn} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}$$

$$\text{So } \rightarrow (1s)^2 (2s, 2p)^8 (3s, 3p)^8 (3d)^{10} (4s)^2$$

→ For  $4s \bar{e}$

$$\sigma = (1 \times 0.35) + (18 \times 0.85) + (10 \times 1) = 25.65$$

$$\text{So } Z^* = 30 - 25.65 = 4.35$$

→ For  $3d \bar{e}$

$$\sigma = (9 \times 0.35) + (18 \times 1) = 21.15$$

$$Z^* = 30 - 21.15 = 8.85$$

$4s \bar{e}$  experiences lower  $Z^*$  so when Zn ionizes it will lose  $4s$  electrons first

### Question (25)

(25). Using Slater rule determine  $Z^*$  for  $4s$  and  $3d \bar{e}$  in Fe ( $Z=26$ ) and Comment on your answer? (Exam Jan 2013)

Answer

$$\text{Fe} \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^6$$

$$\rightarrow (1s)^2 (2s, 2p)^8 (3s, 3p)^8 (3d)^6 (4s)^2$$

$$\text{For } 4s \bar{e} \sigma = (1 \times 0.35) + (14 \times 0.85) + (10 \times 1) = 22.25$$

$$Z^* = 26 - 22.25 = [3.75]$$

$$\text{For } 3d \bar{e} \sigma = (5 \times 0.35) + (18 \times 1) = 19.75$$

$$Z^* = 26 - 19.75 = [6.25]$$

Comment:  $4s \bar{e}$  experience lower  $Z^*$  than  $3d$  so that when Fe ionize to yield  $\text{Fe}^{+2}$  it will lose  $4s$  electrons first

### Question (26)

(26) + write a scientific term for the following:

The portion of total Nuclear charge that a given electron in an atom experiences

Answer

→  $Z_{\text{eff}}$  (effective Nuclear charge  $Z^*$ )

### Question (27)

Choose:  $Z^*$  for  $3d$  electron in Cu is

a) 7.85 (✓) b) 13.05 c) 18.85 (Make it by yourself)



Question (28)  
 28) using Slater's Rule - Calculate  $Z^*$  for Al,  $Al^+$ ,  $Al^{+2}$ ,  $Al^{+3}$ , Discuss the results relative to the expected ionization energies for these species ( $Z=13$ )  
 (exam June 2010)

(Answer)

→ Al electronic Configuration →  $1s^2 2s^2 2p^6 3s^2 3p^1$

→ Putting into groups →  $(1s)^2 (2s, 2p)^8 (3s, 3p)^3$

→  $Z^*$  for  $3s^2$  =  $Z - \sigma = 13 - (2 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 9.5$

$Z^*$  for  $3p$  of Al =  $13 - 9.5 = \boxed{3.5}$

① →  $Al^+ \rightarrow 1s^2 2s^2 2p^6 3s^2$

② →  $(1s)^2 (2s, 2p)^8 (3s)^2$

$\sigma = (1 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 9.15$

$Z^* = 13 - 9.15 = \boxed{3.85}$

$Al^{+2} \rightarrow 1s^2 2s^2 2p^6 3s^1$

③ →  $(1s)^2 (2s, 2p)^8 (3s)^1$

④ →  $\sigma = 1 \times 0.35 + (8 \times 0.85) + (2 \times 1) = 8.8$

⑤ →  $Z^* = 13 - 8.8 = \boxed{4.2}$

$Al^{+3} \rightarrow 1s^2 2s^2 2p^6$

→  $(1s)^2 (2s, 2p)^8$

$\sigma = 1 \times 0.35 + (2 \times 0.85) = 4.15$

$Z^* = 13 - 4.15 = \boxed{8.85}$

	Al	$Al^+$	$Al^{+2}$	$Al^{+3}$
$Z^*$	3.5	3.85	4.2	8.85

Comment: A larger  $Z^*$  on going from Al to  $Al^{+3}$  indicates that there is a stronger attractive force between an outer electron and positive core (nucleus), with this in mind the IE should increase with  $Z^*$  increasing so that values of  $Z^*$  is consistent with IE.

## \* Atomic Parameters \*

1. Atomic Radius    2. Ionization Energy    3. electron affinity  
4. electronegativity.

### Question (29)

29) which has larger radius a) S or  $S^{2-}$  b) Cl or  $Cl^-$  c) Na or  $Na^+$  ✓

1)  $S^{2-} > S$  because:  $S^{2-}$  is the negative ion of S (Anion is larger in its Radius than neutral atom because of greater repulsion that occurs in  $S^{2-}$  than S due to additional two electrons in  $S^{2-}$  so  $S^{2-} > S$  as a result of greater repulsion  $S^{2-}$

(Or Another answer):

\* The nuclear charge is the same for both (S,  $S^{2-}$ ), but the eight valence electrons in  $S^{2-}$  experience a greater amount of repulsion (electron-electron) than do six valence electrons in S atom. This extra repulsion in  $S^{2-}$  increases the average distance between valence electrons so electron cloud around  $S^{2-}$  has larger radius.

2)  $Cl^- > Cl$  ~ the same answer as sulfur  $S^{2-} > S$   
1.81 Å 0.99 Å

2.) Na,  $Na^+$  ~  $Na^+ < Na$  :  $Na^+$  is the positive ion (Cation) of Na (Cation isn't larger in its Radius than neutral atom (is smaller) because the reduction in electron-electron repulsion that arises from loss of electrons as cation formation. This reduction in repulsion makes the nucleus have a stronger pull on the remaining electrons so cation will have smaller size due to attraction of nucleus increasing.

### Question (30)

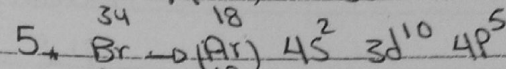
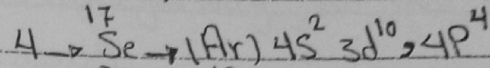
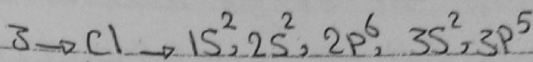
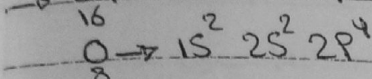
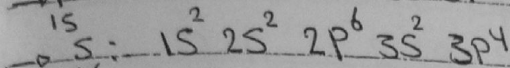
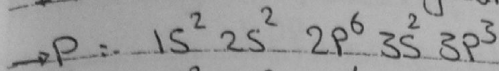
From your knowledge of periodic trend arrange the following elements in order of increasing atomic radius :-

P  $\leq$  S  $\leq$  Br  $\leq$  Cl  $\leq$  Se, O  
15    16    35    17    34    8



## Answer

The electronic Configuration of these elements will be:



Their location in Periodic table as follows

					5A	6A	7A	
Se	Br	P	S	Cl	O	Period 2		
1-17A	1-14A	1-10A	1-6A	0-9A	P <sub>15</sub>	S <sub>16</sub>	Cl <sub>17</sub>	Period 3
					Se <sub>34</sub>	Br <sub>35</sub>		Period 4

①  $\rightarrow$  (Period 2 & group 6A) is the least in atomic number and is expected to be the Smallest in atomic radius due to high  $Z^*$  of Nucleus on electrons (locate at the top of right at Second Period) So it will be the Smallest one.

② For Atoms with the largest atomic number (Se) or (Br) Both are located in fourth period (Groups 6A & 7A) respectively across a period from left to right atomic Radius decreases due to high  $Z^*$  so that Se is the largest followed by Br in atomic radius.

③ Remaining elements are P, S, Cl with atomic number 15, 16, 17 and locates at (Period 3) (Groups 5A, 6A, 7A) respectively across a period atomic Radius decreases on going from left to right

So that Their order will be P > S > Cl

So Final order will be Se > Br > P > S > Cl > O

④ By taking into Consideration  $Z^*$  we obtain the Same order.

## Question (3.1)

③1 Use the electronegativity values to arrange the following bond in order by increasing Polarity: a) P-H < O-H < C-Cl

b) Co < C-S < H-Br where values of electronegativities will be

H = 2.1 < O = 3.5 < C = 2.5 < Cl = 3 < S = 2.5 < Br = 2.8

\* Answer \*

\* The largest Polarity of bond will be to the greater value in difference of electronegativity so that in

$$(1) \quad P-H = 2.19 - 2.1 = 0.09$$

$$O-H = 3.5 - 2.1 = 1.4$$

$$C-Cl = 3 - 2.5 = 0.5$$

So order of increasing polarity will be  $P-H < C-Cl < O-H$

$$(2) \quad C-O = 3.5 - 2.5 = 1$$

$$C-S = 2.5 - 2.5 = 0$$

$$H-Br = 2.8 - 2.1 = 0.7$$

So order of increasing polarity will be  $CS < H-Br < C-O$

Question (32) \*

\* Arrange the following ions in order of decreasing ionic radius

$$a) \quad Mg^{+2} < F^{-} < O^{-2} \quad (Z_{of} \quad Mg=12 < O=8 < F=9)$$

$$b) \quad Cl^{-} < Ca^{+2} < P^{-3} \quad (Z_{of} \quad Cl=17 < Ca=20 < P=15)$$

Answer

(a) order of decreasing will be  $O^{-2} > F^{-} > Mg^{+2}$

\*  $Mg^{+2}$  will be the smallest of them in atomic radii due to loss of  $2e^{-}$  (valence  $e^{-}$ ) and Repulsion - Repulsion of electrons will be reduced and  $I^{+}$  will be the greatest so that  $Mg^{+2}$  will be the smallest

\*  $O^{-2} < F^{-} \rightarrow O^{-2}$  will be the greatest and  $> F^{-}$  due to increasing  $e^{-}$  Repulsion - Repulsion Reduction and decreasing  $I^{+}$  therefore atoms will be larger so  $O^{-2} > F^{-}$

Final order will be  $O^{-2} > F^{-} > Mg^{+2}$

Note that All  $O^{-2}, F^{-}, Mg^{+2}$  are isoelectronic (have the same electronic Configuration)



b)  $\text{Ca}^{+2} < \text{Cl}^- < \text{P}^{-3}$   $\rightarrow$  the order will be  $\text{P}^{-3} > \text{Cl}^- > \text{Ca}^{+2}$

①. Both  $\text{P}^{-3} < \text{Cl}^- < \text{Ca}^{+2}$  are isoelectronic (have the same electronic configuration of Ar(18) in its valence shell)

②.  $\text{Ca}^{+2}$  will be the smallest of them in atomic radius due to loss of  $2e^-$  (valence  $e^-$ ) and Repulsion - Repulsion of electrons will be reduced and  $Z^*$  will be the largest so that  $\text{Ca}^{+2}$  will be the smallest ion

③.  $\text{P}^{-3} < \text{Cl}^-$  (locate in the same period)  $\text{P}^{-3}$  will be the greatest  $> \text{Cl}^-$  due to increasing Repulsion - Repulsion of electrons and decreasing  $Z^*$  so  $\text{P}^{-3}$  will be larger than  $\text{Cl}^-$

Final order will be  $\text{P}^{-3} > \text{Cl}^- > \text{Ca}^{+2}$

Note that you can use  $Z^*$  values to solve this question

### Question (33)

③. What group in the periodic table has elements with the most electronegativities for each period? Determine the element that has the largest electronegativity? and electron affinity?

Answer

\* The group that has elements with the most electronegativity and electron affinity is the group 7A  $\rightarrow$  group of Halogens

\* The elements of high electronegativity & electron affinity:

1) for electronegativity F is the highest one as electronegativity increases across a period and decreases across a group from top to bottom.  $\text{F} \rightarrow$  the top right in periodic table due to small radius and high  $Z^*$  so it is the highest in electronegativity.

\* The elements of highest electron affinity is Cl. Although it is expected to be F as electron affinity increases across a period from left to right as electronegativity and decreases across a group from top to bottom but in F the entering  $e^-$  will suffer from the repulsion of nine electrons of F (small size) but in Cl repulsion is less so the highest element in electron affinity is Cl.

### Question (34)

34. From your knowledge of periodic trends arrange the following elements in order of increasing ionization energy  $\text{Al}$ ,  $\text{Ga}$ ,  $\text{Si}$   
13 31 14

Answer

\* electronic Configuration of these elements is:-

①.  $\text{Al}$ :  $[\text{Ne}] 3s^2 3p^1$  Period 3, Group 3A ③  $\text{Si}$ :  $[\text{Ne}] 3s^2 3p^2$   
②.  $\text{Ga}$   $[\text{Ar}] 4s^2 3d^{10} 4p^1$  Period 4, Group 3A Period 3, Group 4A

3A		4A
Al	Period 3	Si
Ga	Period 4	

①  $\text{Al}$  &  $\text{Si}$  locate in the same period (third) this means that  $\text{Al} > \text{Si}$  in atomic radius and therefore Ionization energy of  $\text{Si} > \text{Al}$ .  
 $\text{Si (IE)} > \text{Al (IE)}$

②  $\text{Al}$  &  $\text{Ga}$  locates in the same group so  $\text{Ga}$  has bigger size and thus smaller IE than  $\text{Al}$

So  $\text{IE Al} < \text{IE of Si}$  and  $\text{IE Ga} < \text{IE of Al}$

Final order will be  $\text{IE of Si} > \text{Al} > \text{Ga}$ .

### Question (35)

35. Describe the general trend in electronegativity of the elements in the periodic table both across a period and down a column?

Answer

\* electronegativity increases across a period from left to right due to the stronger attraction that the atoms have as the nuclear charge increases as atomic radius decreases and electrons still near to nucleus

\* electronegativity decreases on moving down a group due to longer distance between nucleus and valence electron shell (large atomic radius) so atoms has low ability to attract electrons.

Because of large atomic Radius

**F** → the highest element in electronegativity.



### Question (36)

Describe the trends shown by the radii of monoatomic atoms for main group elements (s & p block) both across a period and down a column?

### Answer

- Atomic radius increases on moving down a group because:
  - Valence electrons are found in orbitals of successively high principal quantum number and occupy larger orbitals. So atomic radius increases.
- Atomic radius decreases across a period on moving from left to right because valence electrons enter orbitals of the same shell. Increase in  $Z^+$  across a period draws  $e^-$  and results in a decrease in atomic radius.

### Question (37)

(37). Arrange each of the following pairs in order of increasing radius and explain the order a) Rb,  $Rb^+$  b) Se,  $Se^{2-}$

c) Ca,  $Ca^{+2}$

### Answer

1)  $Rb^+ < Rb$       2)  $Se^{2-} > Se$       3)  $Ca^{+2} < Ca$

Answer in the same way as Question No (29)

### Question (38)

(38). Arrange the following by increasing ionic radius

$Se^{2-}$ ,  $Te^{2-}$ ,  $S^{2-}$  Explain how arrived at this order

$Se \rightarrow Z = 34$        $Te \rightarrow Z = 52$        $S \rightarrow Z = 16$

### Answer

The electron configuration of these ions are: 1)  $Se^{2-} (Ar) 4s^2 3d^{10}, 4p^6$

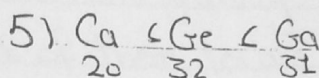
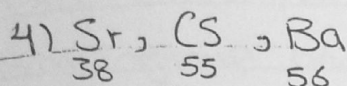
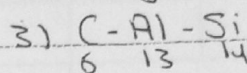
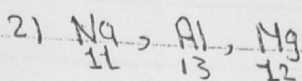
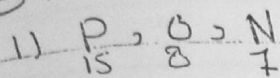
2)  $Te^{2-} \rightarrow [Kr] 5s^2, 4d^{10}, 5p^6$  3)  $S^{2-} \rightarrow Ne, 3s^2 3p^6$

The order will be  $Te^{2-} > Se^{2-} > S^{2-}$  because the atomic number of  $Te^{2-} > Se^{2-} > S^{2-}$  and they locate in period (5, 4, 3) down a group (the same group) from neutral atoms.

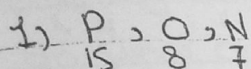
S	3
Se	4
Te	5

### Question (39)

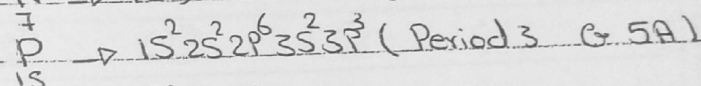
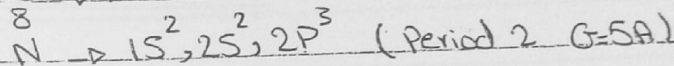
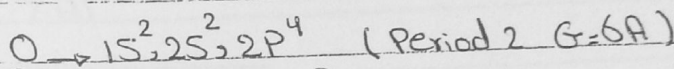
39) Arrange the following in order of increasing electronegativity.



### (Answers)



a) electronic Configuration :



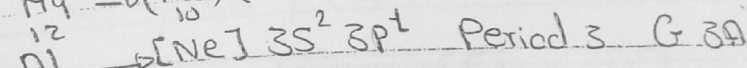
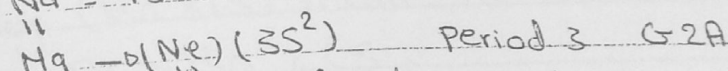
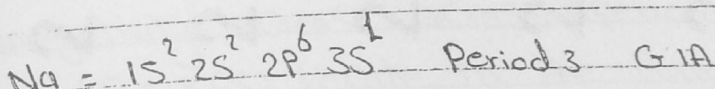
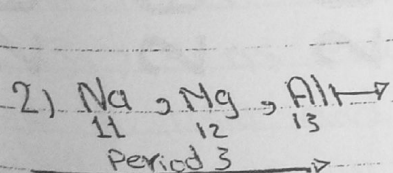
their location will be

5A	6A	
N	O	Period 2
P		Period 3

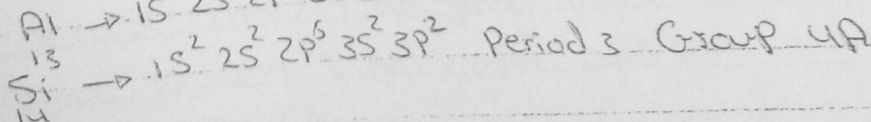
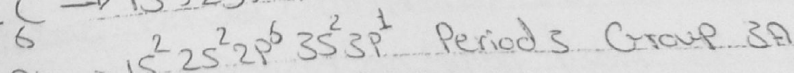
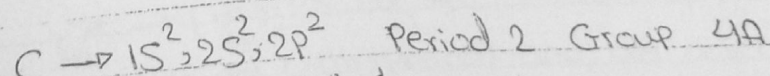
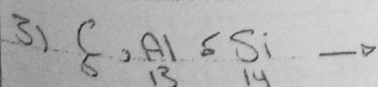
$\rightarrow N, O$  locate in Period 2, So electronegativity of  $O > N$  as it increases across a period

$\rightarrow N, P$  locate in group 5A So electronegativity of  $N > P$  as it decreases down a group

Final order will be  $O > N > P$



All of them locate in the same Period and as electronegativity increases across a period so because of Atomic number final order will be  $Al > Mg > Na$





→ their location will be

3A	4A	
	C	Period 2
Al	Si	Period 3

• C, Si → C > Si as electronegativity decreases on moving down a group and both of them locate at the same group (4A)  
 • Al, Si → Si > Al as they are at the same period (3rd) and electronegativity increases across a period  
 So final order is C > Si > Al

4)  $\overset{38}{\text{Sr}}, \overset{55}{\text{Cs}}, \overset{56}{\text{Ba}}$

• electronic Configuration is:  
 1)  $\overset{38}{\text{Sr}} : [\text{Kr}] 5s^2$  group 2A, Period 5  
 2)  $\overset{55}{\text{Cs}} : [\text{Xe}] 6s^1$  group 1A, Period 6  
 3)  $\overset{56}{\text{Ba}} : [\text{Xe}] 6s^2$  group 2A, Period 6

their location will be

1A	2A	
	Sr	Period 5
Cs	Ba	Period 6

• Ba < Sr as electronegativity decreases on going down a column  
 So Sr > Ba as they locate in the same group (2A)  
 • Ba > Cs as electronegativity increases across a period from left to right and they (Ba & Cs) locate in the same period  
 So final order will be Sr > Ba > Cs

5)  $\overset{20}{\text{Ca}}, \overset{31}{\text{Ga}}, \overset{32}{\text{Ge}}$  : electronic Configuration is:

$\overset{20}{\text{Ca}} \rightarrow [\text{Ar}] 4s^2$  Period 4 / G 2A  
 $\overset{31}{\text{Ga}} \rightarrow (\text{Ar}) 4s^2 3d^{10} 4p^1$  Period 4 / G 3A  
 $\overset{32}{\text{Ge}} \rightarrow (\text{Ar}) 4s^2 3d^{10} 4p^2$  Period 4 / G 4A

• They locate in the same Period (4th period) and as electronegativity increases across a period on going from left to right  
 So because of Atomic Number Final order will be  
Ge > Ga > Ca

### Question (40)

40. Give an explanation for the following:

a) The electron Configuration of Cr is  $[\text{Ar}] 4s^1 3d^5$  rather than  $[\text{Ar}] 4s^2 3d^4$

b) The electron Configuration of Ti is  $[\text{Ar}] 4s^2 3d^2$  but  $\text{Cr}^{+2}$  is  $[\text{Ar}] 3d^4$

### Answer

① Cr is  $\rightarrow [\text{Ar}] 4s^1 3d^5$  because lower total energy may be obtained by forming half-filled d-subshell as stability of element increases when d-subshell is half-filled or completely filled.

② Ti is  $[\text{Ar}] 4s^2 3d^2$  as Ti is a transition element so its valence shell involves filling of electrons in d-subshell. because 4s level is lower in energy than 3d so we fill 4s electrons first then 3d.

\*  $\text{Cr}^{+2} \rightarrow (\text{Ar}) 3d^4$  losing  $2e^-$  from Cr and forming  $\text{Cr}^{+2}$  will take place by removing  $2e^-$  of 4s first as it is lower in energy so we can't need high E for removing 4s electrons. Energy required for removing e from 4s < 3d due to lower  $Z^*$  on 4s electrons.

### Question (41)

41. Select the better choice in each of the following and explain your answer?

1) Higher IE  $\therefore$  ① Ca or Ga  
20 31  
 Ionization energy

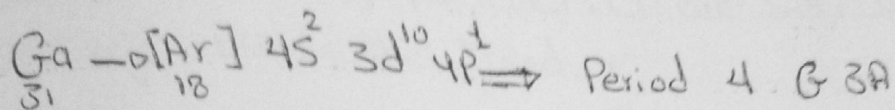
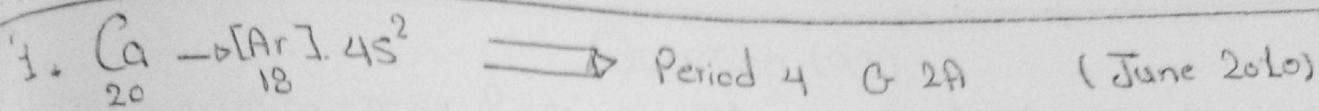
② Mg or Ca  
12 20

2) Higher electron affinity Si or P ✓  
14 15

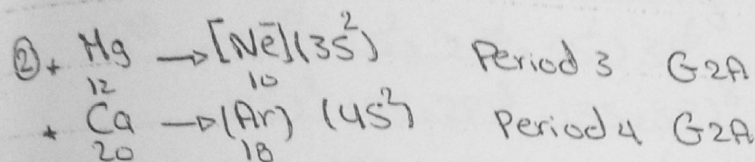
3) More likely configuration for  $\text{Mn}^{+2}$  :  $[\text{Ar}] 4s^2 3d^5$  or  $[\text{Ar}] 3d^5$

### Answer



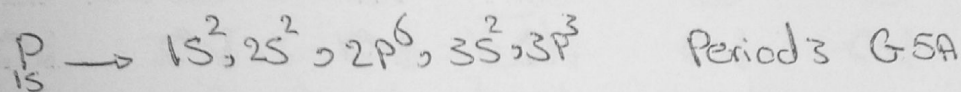


① locate in the Same Period and ionization Energy increases across a period on moving from left to right so ~~Ga~~  $\text{Ca} > \text{Ga}$  in IE  $\text{Ca} > \text{Ga}$



2A	
Mg	3
Ca	4

locate in the Same group & IE decreases on going down a group  
 So that  $\text{Mg} > \text{Ca}$  in IE



due to electron Configuration they locate in the Same period (3rd Period) and IE increases across a period on going from left to right and Electron affinity also so  $\text{P} > \text{Si}$  in electron affinity but in this case  $\text{Si} > \text{P}$  because an electron added to p must go into orbital that already contains an electron so there will be Repulsion between added electron and original one and less energy is Released during this addition so  $\text{Si} > \text{P}$  in electron affinity.

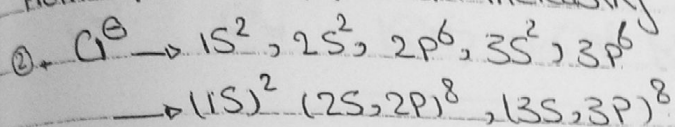
4) More likely electron Configuration of  $\text{Mn}^{2+}$  is  $[\text{Ar}] 3d^5$  because when Mn ionizes it loses less tightly bound electrons to the nucleus ( $4s$ ) according to  $1^{\circ}$  and also  $4s < 3d$  in energy so we don't need high energy for removing  $4s$  electrons as energy Required for removing  $3d$  electrons.

### Question (42)

The ionization energy for  $\text{Cl}^-$ ,  $\text{Cl}$ ,  $\text{Cl}^+$  are 849, 1251 & 2300 kJ/mol respectively. Explain this trend. ✓

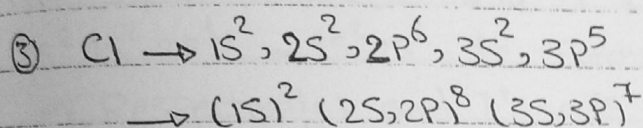
### Answer

① Ionization energy: Minimum energy required to remove the least tightly bound electron to the nucleus. It increases across a period from left to right with increasing  $Z^*$  & decreasing Radius.



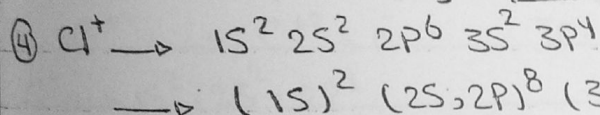
$$\sigma = (7 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 11.25$$

$${}_{3p} Z_{\text{eff}} = 17 - 11.25 = \boxed{5.75}$$



$$\sigma = (6 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 10.9$$

$$Z^* = 17 - 10.9 = \boxed{6.1}$$



$$\sigma = (5 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 10.55$$

$$Z^* = 17 - 10.55 = \boxed{6.45}$$

$\text{Cl}^-$	$\text{Cl}$	$\text{Cl}^+$	$\rightarrow Z^*$
5.75	6.1	6.45	

Larger  $Z^*$  on moving from  $\text{Cl}^-$  to  $\text{Cl}$  indicate that valence electron in  $\text{Cl}^+$  experience highest  $Z^*$  so that removing of electron from  $\text{Cl}^+$  will need highest energy = 2300 kJ/mol then  $e^-$  in  $\text{Cl}$  which experience  $Z^* = 6.1$  so will need lower energy, the least one is  $\text{Cl}^-$  due to increase in atomic radius & decrease in  $Z^*$ . So will require the least amount of energy.

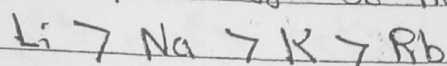
### Question (43)

⑤ Why are the ionization energies of alkali metals in the order  $\text{Li} > \text{Na} > \text{K} > \text{Rb}$



## Answer

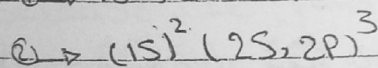
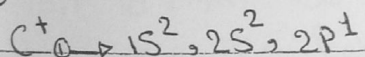
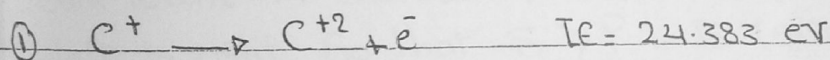
All these elements locate in the same group (1A) and on downing a group IE decreases because atomic radius increases and Removal of electrons doesn't require higher values of energy because valence electrons on downing a group are less held by the nucleus and loss of electrons doesn't require higher values of energy so the trend will be



## Question (44)

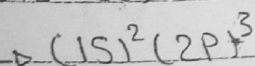
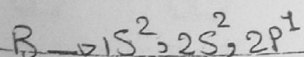
44. The Second ionization energy of  $\text{C}^+ \rightarrow \text{C}^{2+} + e^-$  and First ionization energy of boron ( $\text{B} \rightarrow \text{B}^+ + e^-$ ) Both fit the reaction  $1s^2 2s^2 2p^1 \rightarrow 1s^2 2s^2 + 1e^-$  Compare the two IE 24.383 eV and 8.298 eV respectively and  $7^+$  is it adequate explanation of difference in IE? if not suggest other factor.

## ANSWER



③  $\rightarrow \sigma = (2 \times 0.35) + (2 \times 0.85) = 2.4$

$7^+ = 6 - 2.4 = \boxed{3.6}$



$\sigma = (2 \times 0.35) + (2 \times 0.85) = 2.4$

$7^+ = 5 - 2.4 = \boxed{2.6}$

	B	C <sup>+</sup>
z <sub>eff</sub>	2.6	3.6

larger  $7^+$  on  $\text{C}^+$  indicate that valence  $e^-$  is drawn closer to the nucleus and therefore require higher E for removal than B needs

AS atoms:-

IE of C > IE of B

and IE C > IE of B as they locate in the same period

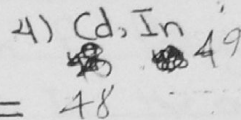
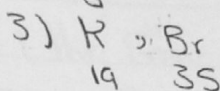
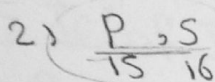
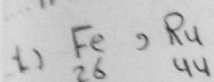
Second IE C >> First IE C > IE of B

So  $\text{C}^+ \gg \text{B}$

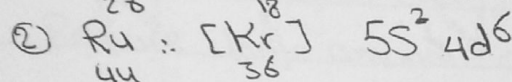
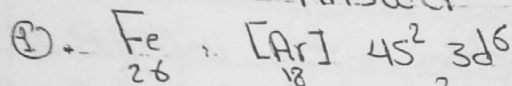
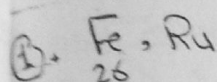
IE	IE
24.383	8.2 eV

# Question (45)

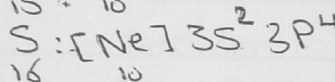
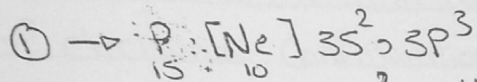
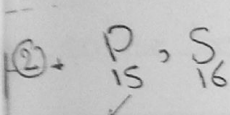
⑤. In each of the following pairs pick up the element with the higher IE and explain your choice?



## Answer



Both elements locate in group (8) and on going a group (up to down) IE decreases and Ru is down Fe so that Fe > Ru in IE



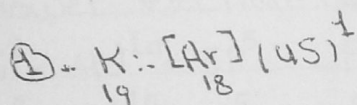
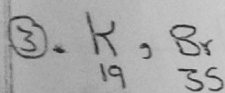
3p <sub>x</sub>	3p <sub>y</sub>	3p <sub>z</sub>
↑	↑	↑

Period 3 G 5A

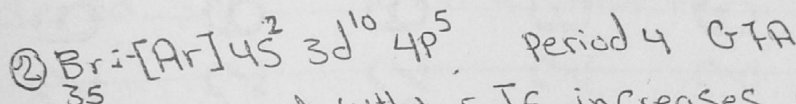
3p <sub>x</sub>	3p <sub>y</sub>	3p <sub>z</sub>
↑↓	↑	↑

Period 3 G 6A

Both locate in the same period so IE increases across a period from left to right so that S > P in IE (expected) but here P <sub>15</sub> > S <sub>16</sub> because IE falls to lower value in S as fourth electron is paired inside 3p orbital so that it's repulsion doesn't require higher energy (Minimize Repulsion Force)

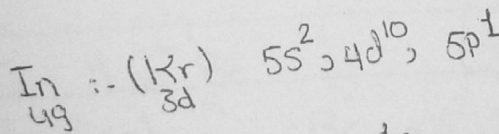
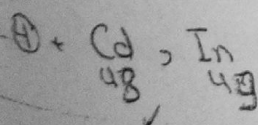


Period 4 G 1A

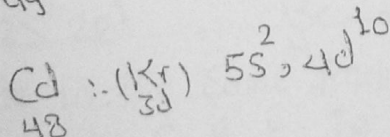


period 4 G 7A

Both locates in the same period (4th) & IE increases across a period from left to right so IE Br > IE K #.



Period 5 G 3A



Period 5 G 2A



Both locate in the same period (5th) and IE increases across a period so  $IE_{In} > IE_{Cd}$  (expected) but here

$IE_{Cd} > IE_{In}$  (the same reason for falling IE between Be and B) Because on going from Cd to In the outermost electron occupies 5p electron and hence is less strongly bound and opening 5s subshell requires energy so removal of e in this case doesn't require large amount of energy.

5)  $C \rightarrow N$   
6 7

$IE_N > IE_C$  why?

Question (46)

Q. on the basis of electronic configuration explain why?

Answer

a) Sulfur has lower electron affinity than Chlorine (Jan 2013)

$S \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^4$  Period 3 G 6A

$Cl \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^5$  Period 3 G 7A

They locate in the same period (3d period) and electron affinity increases across a period from left to right so that  $S < Cl$ .

b) Iodine has lower electron affinity than bromine

$I \rightarrow [Kr] 5s^2, 4d^{10}, 5p^5$  Period 5 Group 7A

$Br \rightarrow [Ar] 4s^2, 3d^{10}, 4p^5$  Period 4 Group 7A

Both elements locate in the same group (7A) and electron affinity decreases on going down a group so that  $I < Br$  in electron affinity.

c) B has lower IE than Be (Jan 2013)

$B \rightarrow 1s^2 2s^2 2p^1$

$Be \rightarrow 1s^2 2s^2$

Both elements locate in the same period and IE increases across a period from left to right so that  $IE_B > IE_{Be}$  (expected)

but in fact  $IE_{Be} > IE_B$  because in B the outermost electron occupies a 2p orbital and hence is less tightly bound to the nucleus so value of IE falls back to lower value in B

4) Chlorine has lower IE than Fluorine?

Cl  $\rightarrow 1s^2 2s^2 2p^6 3s^2 3p^5 \rightarrow$  Period 3 G 7A

${}^{17}_9F \rightarrow 1s^2 2s^2 2p^5 \rightarrow$  Period 2 G 7A

Both elements locate at the same group (7A) and IE on moving down a group decreases from top to bottom so that  $IE_{Cl} < IE_{F}$

### Question (47)

Q. The size of transition metal atoms decreases slightly from left to right in the periodic table? why does the decrease in size occur at all? and why it is so gradual?

(Answer)

It decreases at all across a period on moving from left to right because across a period the valence electrons enter orbitals of the same shell, the increase in  $Z^*$  across a period draws in the electrons and leads to decrease in atomic radius.

2) decrease is so gradual because in transition metal we fill (d) orbital successively by electrons and the added electron when enters d orbital feels great repulsion with other electrons (2 electrons in d orbital repel each other with higher energy than 2 electrons in s or p subshell).

This effect makes nucleus effect on valence electrons less than expected so atomic radius decreases gradually.

Another factor mass of electron in elements of high atomic number.



Question 148)

48) Predict the largest and smallest in each series (June 2010)

①  $\text{Se}^{-2}$ ,  $\text{Br}^{-}$ ,  $\text{Rb}^{+}$ ,  $\text{Sr}^{+2}$       ③  $\text{Co}$ ,  $\text{Co}^{+2}$ ,  $\text{Co}^{+3}$ ,  $\text{Co}^{+4}$

②  $\text{Y}^{+3}$ ,  $\text{Zr}^{+4}$ ,  $\text{Nb}^{+5}$

Answer

1)  $\text{Se}^{-2}$ ,  $\text{Br}^{-}$ ,  $\text{Rb}^{+}$ ,  $\text{Sr}^{+2}$  all are isoelectronic and have the same electronic Configuration of  $[\text{Kr}] + 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6$

So  $d$  for valence  $e$  For all atoms  $^{36}$  will be the same and ionic radii decreases with increasing magnitude of Nuclear Charge (Number of Protons)

$\text{Se}^{-2} = 34$  Proton       $\text{Rb}^{+} = 37$  Proton

$\text{Br}^{-} = 35$  Proton       $\text{Sr}^{+2} = 38$  Proton

Ionic Radius decreases in the order  $\text{Se}^{-2} > \text{Br}^{-} > \text{Rb}^{+} > \text{Sr}^{+2}$

2)  $\text{Y}^{+3}$ ,  $\text{Zr}^{+4}$ ,  $\text{Nb}^{+5}$

+ all are isoelectronic and have the same electronic Configuration of  $[\text{Kr}] + 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6$

So ionic radius decrease with increasing Nuclear Charge (Number of Protons) and  $z^{+}$

For  $\text{Y}^{+3} = 39$  Proton

$\text{Zr}^{+4} = 40$  Proton

$\text{Nb}^{+5} = 41$  Proton

So ionic Radius decreases in the order of  $\text{Y}^{+3} > \text{Zr}^{+4} > \text{Nb}^{+5}$

3)  $\text{Co}$ ,  $\text{Co}^{+2}$ ,  $\text{Co}^{+3}$ ,  $\text{Co}^{+4}$  will be  $\text{Co}^{+2} > \text{Co}^{+3} > \text{Co}^{+4}$

Radius of Positive Cation is always smaller than its Neutral atom because of Reduction of electron Repulsion - Repulsion that Result from loss of electrons due to Cation formation. This reduction in repulsion make nucleus has Stronger Pull on remaining electrons and as the Number of removed electrons increases the atomic Radius decreases so  $\text{Co}^{+4} < \text{Co}^{+3} < \text{Co}^{+2} < \text{Co}$

⇒ Give suitable reason: Second IE  $>$  First IE for any element

Give example (Jan 2013)

f  
Cl  
Br  
I  
increases

Question 149)  
49\* which of the following atoms has the largest diameter? (June 2012)  
a. F      b. Cl      c. Br      d. I (V)

Answer

2\* which element would have the greater difference between the first IE and the second IE?  
(June 2012)

a) K (V)      b. Ca<sub>20</sub>      c. Both of them

Answer

(a) \* K (why)?

Question 150)

50\* write a scientific term for the following:

a) The energy required to remove an electron from the outermost valence orbital of gaseous atom (Ionization energy) (Jan 2013)

b) The power of an atom of the element to attract electrons to itself when it is apart of a compound (electronegativity) (Jan 2013)

(c) Electron affinity of oxygen = -141 kJ/mol means <sup>heat released</sup> -----?  
→ heat Released. (Jan 2013)

(d) \* The element with highest IE is <sup>1s<sup>2</sup> 2s<sup>2</sup> 2p<sup>3</sup></sup> 3) N (V) <sup>1s<sup>2</sup> 2s<sup>2</sup> 2p<sup>4</sup></sup> 4) O (Jan 2010)

a) B<sub>5</sub>

2) C<sub>6</sub>

→ N. why? → ~~half~~ half filled → more stable need more energy to remove an electron than B or C or O

(E) which of the following has the largest atomic radius?  
(Jan 2010)

a) Al<sub>13</sub>

2) Ca<sub>20</sub>

3) F<sub>9</sub>

4) K (V)<sub>19</sub>

5) S<sub>16</sub>

(F) \* which have the smallest radius

1) K<sub>19</sub>

2) Fe<sub>26</sub>

3) As<sub>33</sub>

4) Br<sub>35</sub>

5) Kr (V)<sub>36</sub> (Jan 2010)



G) which have the highest IE (Jan 2010)  
1) O 2) Be 3) F 4) C 5) B  
8 4 9 6 5

H) which have the highest electronegativity? (Jan 2010)  
1) P 2) N 3) K 4) As 5) Be  
3 7 19 33 4

I) The decreasing order of electron affinity F, Cl, Br is  $F > Cl > Br$   
(X) why (Jan 2010)  $Cl > F > Br$

### Question 151

151) define the following (Jan 2010)

- First ionization energy?
- electron affinity?
- Covalent and ionic radius?

B) Fluorine is the more electronegative element but Cs is (V)  
the Smallest electronegative element in the Periodic table (Jan 2010)

C) Positive ions are bigger than their Neutral atom but Negative  
ions are Smaller than their Neutral atom (X) (June 2010)

D) Elements in the group 1A, 2A and 3A don't follow octet rule (X)  
(June 2010)

E) More likely electron Configuration of  $Mn^{+2}$  is  $[Ar] 4s^2 3d^3$  (X)  
why. (June 2010)

F) Define IE and why Na has less IE than Li (June 2010)

G) The outer electronic Configuration of Cu.  $3d^9 4s^2$  (X) why?

H) which of the following have the greatest electronegativity?  
 a) Si b) P c) N d) O (v) why?

E) which have the smallest radius:

1)  $O^{2-}$  2)  $Na^+$  3)  $F^-$  4)  $Al^{3+}$  why.

(8)  $Zn^{+2}$  exhibits Paramagnetism due to loss of electrons from 3d-orbital (x)

$\rightarrow Zn^{+2}$  exhibits diamagnetism, Formation of  $Zn^{+2}$  involves loss of  $2e^-$  from 4s orbital, the resultant cation has  $10e^-$  in 3d orbital and doesn't have unpaired  $e^-$  so it is diamagnetic

(9) \* In group 1A (Alkali metals), IE decreases down a group so Li is a poor reducing agent. (x)

Reducing strength is measured by negative of reduction potential not ionization so Li is the strongest reducing agent among Alkali metals.

### Question (52) \*

(52) \* Ionization energy should depend on  $Z^*$  that hold electrons in atoms, Calculate  $Z^*$  for (N) and (As). Do their IE match with  $Z^*$ ? If not, what other factor influence on IE?  $N \Rightarrow Z=7$   $P \Rightarrow Z=15$   $As \Rightarrow Z=33$  ? (Jan 2012)

Answer

①  $Z^*$  For valence electron of N

1)  $N \rightarrow 1s^2 2s^2 2p^3$

$\rightarrow (1s)^2 (2s, 2p)^5$

$\rightarrow S = (4 \times 0.35) + (2 \times 0.85) = 3.1$

$\rightarrow Z^* = 7 - 3.1 = \boxed{3.9}$

$Z^*$  For valence electron of P

$P \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

$\rightarrow (1s)^2 (2s, 2p)^8 (3s, 3p)^5$

$\rightarrow S = (4 \times 0.35) + (8 \times 0.85) + (2 \times 1) = 10.2$

$\rightarrow Z^* = 15 - 10.2 = \boxed{4.8}$



$AS \rightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^3$   
 $\rightarrow (1s)^2 (2s, 2p)^8 (3s, 3p)^8 (3d)^{10} (4s, 4p)^3$   
 $\sigma = (4 \times 0.35) + (18 \times 0.85) + (10 \times 1) = 26.7$   
 $Z^* = 33 - 26.7 = 6.3$

N	$1s^2, 2s^2, 2p^3$	Period 2	G-5A	3.9	5A
P	$[Ne] 3s^2, 3p^3$	Period 3	G-5A	4.8	N
AS	$[Ar] 4s^2, 3d^{10}, 4p^3$	P-4	G-5A	6.3	P
					AS

$Z^*$  increase on going from N to AS will indicate that IE from N to AS will increase due to small Radius but this is not true as IE of N > IE of AS because on moving down a group IE decreases. So values of  $Z^*$  doesn't match with Ionization energy trend.

(Factors)

size  
 On moving from N to AS we fill valence  $\bar{e}$  each time in new shell (we open a new shell each time) so that distance between nucleus and valence  $\bar{e}$  become larger which make isolation of valence electron each time become easier. So first factor  $\rightarrow$  size of atom which become larger on downing a group.

Shielding effect: (electron in the distance between nucleus and electron of interest), as no. of these electrons increases this will make nucleus effect on valence  $\bar{e}$  less stronger and its removal requires less energy. So IE trend for these elements will be  $N > P > AS$ .

# Final Revision

## CHAPTER [3]

### Question (1)

Use Lewis electron dot symbols to represent the transfer of electrons from Mg to F atoms to form ions with noble gas configuration (Z of Mg = 12 and Z of F = 9)?

(Answer)

The Lewis electron-dot symbols for the atoms are  $\cdot\ddot{\text{F}}\cdot$  and  $\times\text{Mg}\times$ .  
 Mg atom loses two electrons to assume noble gas configuration.  
 F can accept only one electron to fill its valence shell.  
 So that two F atoms must share in the electron transfer.  
 As follows:

$$\begin{array}{l} \text{F} \rightarrow 1s^2 2s^2 2p^5 \rightarrow \text{F}^- \rightarrow 1s^2 2s^2 2p^6 (\text{Ne}) \\ \text{Mg} \rightarrow 1s^2 2s^2 2p^6 3s^2 \rightarrow \text{Mg}^{+2} \rightarrow 1s^2 2s^2 2p^6 [\text{Ne}] \end{array}$$

### Question (2)

Use Lewis electron dot symbols to represent the transfer of electrons from Mg to O atoms to form ions with noble gas configuration?

(Answer)

The Lewis-electron-dot symbols for the atoms are  $\times\text{Mg}\times$  and  $:\ddot{\text{O}}:$ .  
 Mg atom loses two electrons to assume noble gas configuration.  
 O atom can accept two electrons to fill its valence shell.  
 So that one oxygen atom must share in the electron transfer.  
 As follows:

$$\begin{array}{l} \text{Mg} \rightarrow \text{Ne} \\ \text{O}^{+2} \rightarrow \text{Ne} \end{array}$$



Use Lewis electron-dot symbols to represent the transfer of electrons from Na to Cl to form ions with Noble gas configuration?

### Question (3)

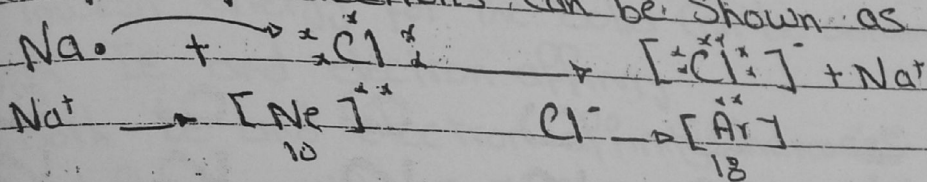
### ANSWER

The Lewis electron-dot symbols for the atoms are  $\cdot\ddot{\text{Cl}}\cdot$  and  $\cdot\text{Mg}\cdot$  ( $\text{Na}\cdot$ )

Na atom loses one electron to assume Noble gas configuration

Cl atom can accept only one electron to fill its valence shell and reach Noble gas configuration

Transfer of electrons can be shown as follows

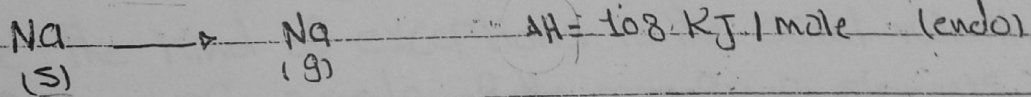


### Question (4)

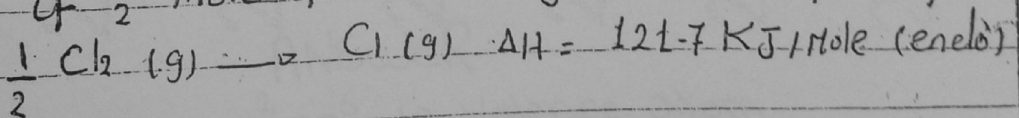
1. Describe the formation of Sodium Chloride crystals from atoms?

### ANSWER

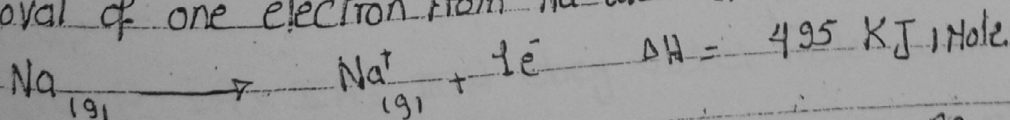
1. Vaporization (Sublimation) of 1 mole of Na(s) to produce Na (vapors) (gaseous state)



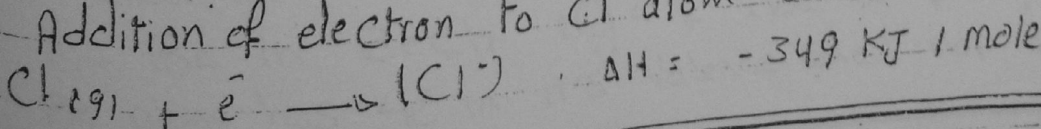
2. Decomposition of  $\frac{1}{2}$  mole of  $\text{Cl}_2$  to give 1 mole of Cl as follows:

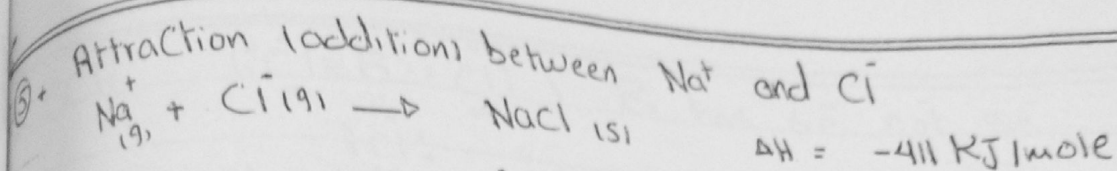


3. Removal of one electron from Na atom (First Ionization energy of Na)



4. Addition of electron to Cl atom (Electron affinity of Cl)





### Question (5)

⑤. Explain what energy terms are involved in formation of an ionic solid from atoms. In what way should these terms change to give the lowest energy possible for the solid?

### ANSWER

Terms of energy are ① Ionization energy ② electron affinity ③ energy of dissociation ④ Energy Released due to formation of solid (NaCl for example) from the combination of  $\text{Na}^+$  with  $\text{Cl}^-$  Ion

→ lowest energy for the solid can be obtained by controlling both IE and electron affinity

→ Both IE and electron affinity increases across a period from left to right as  $Z_{\text{eff}}$  increases and atomic size decreases. So the stable ionic compounds (lattice energy increases) if the IE of metal is small and electron affinity of the non-metals is sufficiently large.

### Question (6)

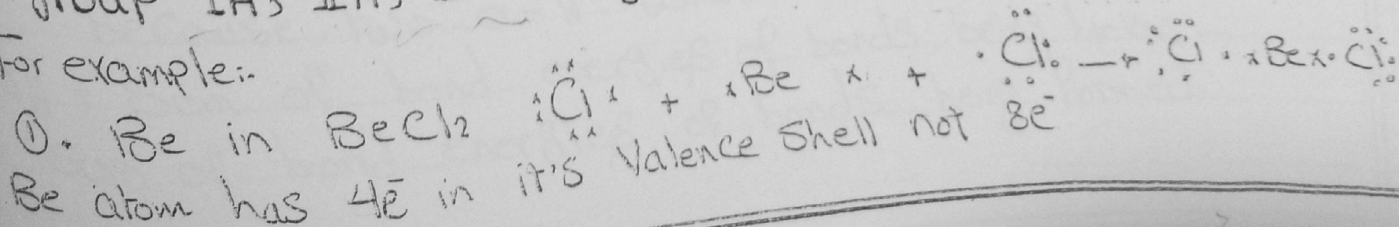
⑤. Describe the kind of exceptions to the octet rule that we encounter in compounds of the main group elements? Give example?

### ANSWER

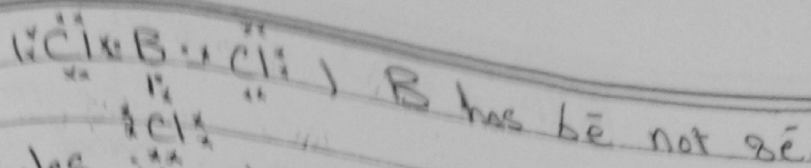
⑤. Octet Rule means  $ns^2 np^6$  in valence atom.

Group - ⑤. Group IA, IIA, IIIA obey octet rule with some exceptions

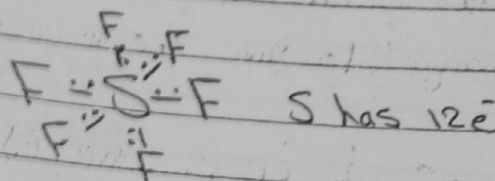
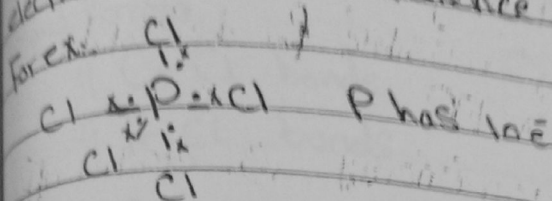
For example:







Many Molecules in which Central atom has more than octet  
 when central atom locates from Period 3 to Period 6  
 due to ability of these atoms to accept more than eight  
 electrons in their valence  $3d$  orbital.



So that Be, B, Period from 3-6 are good exceptions for octet  
 $ns^2 np^6 nd^{10}$  are pseudo noble gas configuration.

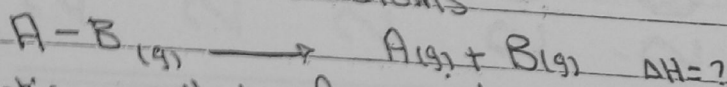
### Question (7)

Define bond energy, Explain how one can use bond  
 energies to estimate the heat of reaction?

### ANSWER

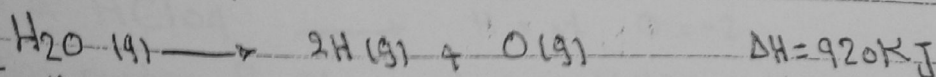
**Bond energy:** It is the heat required to break one mole  
 of Molecules into their individual atoms

For ex:



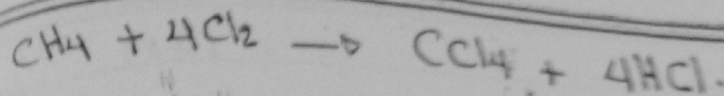
The bond dissociation, enthalpy of molecule derived from  
 Complete atomization.

The mean bond enthalpy  $B$  is the quantity derived by  
 the number of bond dissociated.



In this example the mean  $\text{O}-\text{H}$  bond enthalpy is one-  
 half atomization enthalpy or  $460 \text{ KJ/mole} = \frac{920}{2} = 460 \text{ KJ/mole}$   
 because two  $\text{O}-\text{H}$  bonds dissociated.

$\Delta H = \text{Sum of bond energies of bonds being broken} -$   
 $\text{Sum of bond energies of bonds being formed.}$



ex: bond broken

4 C-H bond

$$= 4 \times 413 = 1652$$

4 Cl-Cl bond

$$= 4 \times 243 = 972$$

$$2624$$

Bond Formed

4 C-Cl bonds

$$4 \times 328 = 1312$$

4 H-Cl bonds

$$4 \times 432 = 1728$$

$$3040$$

$$\Delta H_{\text{Reaction}} = 2624 - 3040 = -416 \text{ KJ/mole} \quad \text{So endo or exo (exothermic)}$$

Question (8)

Consider the molecules  $\text{N}_2\text{H}_4$ ,  $\text{N}_2$  and  $\text{N}_2\text{F}_2$  which molecule has the shortest nitrogen-nitrogen bond and which has the largest?

ANSWER

The Nitrogen-Nitrogen bond should be the shortest in  $\text{N}_2$  where it is the triple bond and the longest in  $\text{N}_2\text{H}_4$  where it is a single bond.

Question (9)

(Oxidation Numbers)

What is the oxidation number of :-

a) + Cl in  $\text{HClO}_4$

$$\text{H} = +1 \quad \text{O} = -2$$

$$\text{Oxidation number of Cl} = +1 + 0 \cdot \text{N}(\text{Cl}) + (-2 \times 4) = 0$$

$$+1 + 0 \cdot \text{N}(\text{Cl}) - 8 = 0$$

$$0 \cdot \text{N}(\text{Cl}) = +8 - 1 = +7$$

b) + S in  $\text{SO}_4^{2-}$

$$0 \cdot \text{N}(\text{S}) + (-2 \times 4) = -2 \quad \therefore 0 \cdot \text{N}(\text{S}) - 8 = -2 \quad \therefore 0 \cdot \text{N}(\text{S}) = 8 - 2 = 6$$



$$\text{S in } \text{SO}_4^{-2} \quad (+6)$$

$$\text{B in } \text{BF}_4^{-1}$$

$$\text{O.N (B)} + (4 \times -1) = -1$$

$$\text{O.N (B)} - 4 = -1$$

$$\text{O.N (B)} = +4 - 1 = +3$$

$$\text{Cr in } \text{CrO}_4^{-2}$$

$$\text{O.N (Cr)} + (4 \times -2) = -2$$

$$\text{O.N (Cr)} - 8 = -2$$

$$\text{O.N (Cr)} = +6$$

$$\text{Al in } \text{Al(OH)}_3$$

$$\text{O.N (Al)} + (3 \times -1) = 0$$

$$\text{O.N Al} - 3 = 0 \quad \text{O.N (Al)} = +3$$

$$\text{E) N in } \text{NO}_2^-$$

$$\text{O.N (N)} + (-2 \times 2) = -1$$

$$\text{O.N (N)} - 4 = -1$$

$$\text{O.N (N)} = +4 - 1 = +3$$

$$\text{G) P in } \text{PO}_4^{-3}$$

$$\text{O.N (P)} + (4 \times -2) = -3$$

$$\text{O.N (P)} - 8 = -3$$

$$\text{O.N (P)} = +8 - 3 = 5$$

### Question (10)

Choose the correct Answer:-

Q. The ion that is isoelectronic with  $\text{CO}$  is:

- a)  $\text{CN}^-$  (✓) b)  $\text{O}_2^+$  c)  $\text{O}_2^-$  d)  $\text{N}_2^+$

Q. which one of the following don't have hydrogen bond?

- a) Phenol b) liquid  $\text{NH}_3$  c) water d) liquid  $\text{HCl}$  (✓)

Q. The bond between two identical non-metal atoms has a pair of electrons ..... Complete? (equally shared between them.)

Q. The molecule having one unpaired electron is

- a)  $\text{NO}$  (✓) b)  $\text{CO}$  c)  $\text{CN}$  d)  $\text{O}_2$

Q. on hybridization of one S and one P orbitals we get:

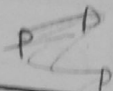
- a) two mutually perpendicular orbitals b) two orbitals at  $180^\circ$  (✓)  
c) Three orbitals in a plane  
d) Four orbitals directed tetrahedrally

- The bonds present in  $N_2O_5$  are ::
- only ionic
  - only Covalent
  - Covalent and Coordinate (✓)
  - Covalent and ionic

### Question (11)

Complete the following Answer:

- There are two  $\pi$  bonds in a nitrogen molecule  $N \equiv N$
- $sp^3$  hybrid orbitals of nitrogen are involved in formation of  $NH_4^+$
- The terminal Carbon atom in butane is  $sp^3$  hybridized
- The valence atomic orbitals of Carbon in Silver acetylide  $Ag-C \equiv C-Ag$  is  $sp$
- In  $P_4O_{10}$  the Number of oxygen atoms bonded to each P atom is Four atoms



### Question (12)

- State whether the following statements are true or False
- Linear overlap of two atomic P-orbitals leads to  $\sigma$  bond (✓)
- $sp^2$  hybrid orbitals have equal S and P character (x)  
 $\frac{1}{3}$  S character and  $\frac{2}{3}$  P character
- The H-N-H bond angle in  $NH_3$  is greater than H-Al-H bond in  $AlH_3$  (✓)  $\rightarrow$  Repulsion is more in  $NH_3$  which make bond angle greater (Exam June 2013)

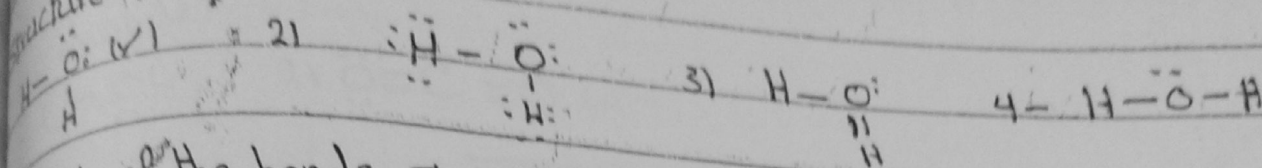
### Question (13)

Choose the correct Answer?

- Flourine atoms tend to ---- ? when they form chemical compounds with metals: 1) lose electrons 2) gain electrons (✓)
- which of the following represent correct formula for Aluminium oxide: 1)  $AlO$  2)  $Al_2O_3$  (✓) 3)  $AlO_2$  4)  $Al_2O$
- What is the formal charge of oxygen atoms of the following  $H-\ddot{O}-H$   $\rightarrow$  a) +3 b) +1 (✓) c) -2



Which of the following represent the correct electron dot structure for water?



Which of the bonds shown by the dash has the greatest polarity?

1)  $\text{H}-\text{Cl}$  2)  $\text{H}-\text{NH}_2$  3)  $\text{H}-\text{O}-\text{H}(\text{V})$  4)  $\text{H}-\text{S}-\text{H}$

Where the electronegativity values for these atoms are:

$\text{H} \rightarrow 2.1$   $\text{N} \rightarrow 3.1$   $\text{O} \rightarrow 3.5$   $\text{Cl} \rightarrow 2.5$   $\text{S} \rightarrow 2.4$  (why  $\text{H}-\text{O}-\text{H}$ )

In which of the following compounds does ionic compound occur?

a)  $\text{NH}_4\text{Cl}$  b)  $\text{CO}_2$  c)  $\text{CH}_4$  d)  $\text{LiBr}(\text{V})$

Which of the following molecules contains only one non-bonding pair of valence electrons?

1)  $\text{NH}_4^+$  2)  $\text{H}-\text{C}\equiv\text{N}$  3)  $\text{C}_2\text{H}_4$  4)  $\text{N}_2(\text{V})$

Which of the following gases would be most soluble in water?

1)  $\text{N}_2$  2)  $\text{NH}_3(\text{V})$  3)  $\text{CH}_4$  4)  $\text{CO}_2$

When a double bond is formed between two atoms, one of the bonds is a sigma bond and the other is pi bond. The pi bond is created by the overlap of:

1)  $\text{sp}^3$  hybrid orbitals 2)  $\text{sp}^2$  hybrid orbital  
3) p orbitals(V) 4) s orbitals

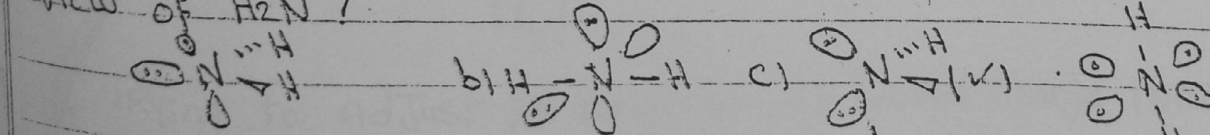
The central atom in  $\text{BrF}_5$  has how many bonding pairs of electrons and non bonding pairs of electrons?

a) (1...5) b) 0...5 c) 5...1(V) d) 5...0

What is the hybridization in  $\text{SF}_4$ ?

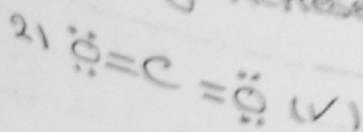
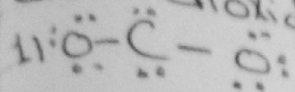
a)  $\text{sp}^2$  b)  $\text{sp}^3\text{d}(\text{V})$  c)  $\text{sp}^3\text{d}^2$  4)  $\text{sp}^3$

Which of the following best represents the 3-dimensional view of  $\text{H}_2\text{N}^-$ ?



The energy of electron in a 4s orbital is less than that in 4s-orbital (V)

14) The Carbon atom in benzene ring is .... hybridized? ( $sp^2$ )  
 15) which of the following represents electron-dot formula of Carbon dioxide?



Question 14)

Lewis Structure & formal charge.

2)  $CCl_2F_2$

3)  $COCl_2$  4)  $CO_2$  5)  $SO_3^{2-}$

(ANSWER)

1 General Steps

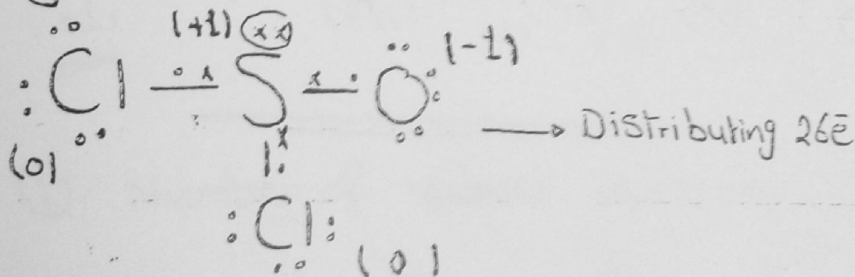
1. For Drawing Lewis Structure correctly we must follow these rules:-
1. Decide Central atom correctly
2. Count all valence electrons of the atom in each example
3. If there is anion add one electron, If cation subtract one electron
4. Place out pair of electrons in each bond.
5. Complete the octets of atoms bonded to the Central atom.
6. Place any additional electrons on Central atom.
7. If Central atom has less than octet, form multiple bonds.

1.  $SOCl_2$  ..

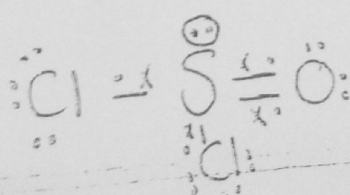
Number of valence electrons =  $S(6) + O(6) + Cl(7 \times 2) = 26e^-$

2.  $\rightarrow$  More electropositive will be Central atom with O & Cl bonded to it

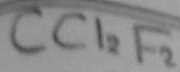
3. Distributing the electrons.



4. Trying to Adjust All atoms on Zero formal charge (Neutral compound) by forming double bond with oxygen

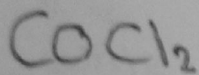
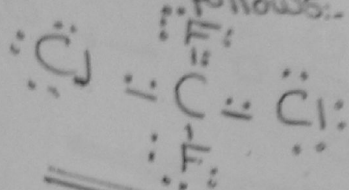






No of valence  $e^- =$

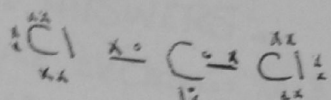
Distributing  $32e^-$  as follows:-



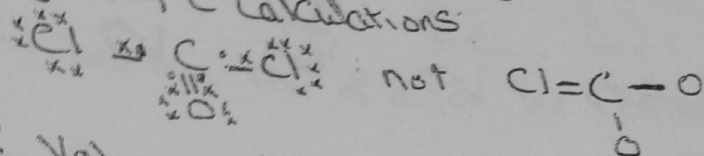
① C most electropositive is the central atom

② Number of valence electrons =

Distributing 24 electrons

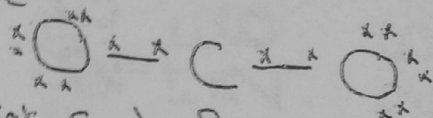


C atom still less than octet so that by forming double bond with oxygen according to F-C calculations

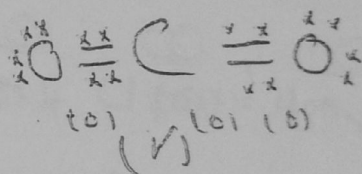
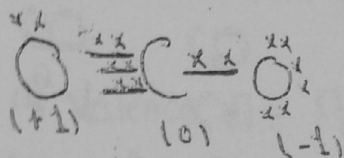
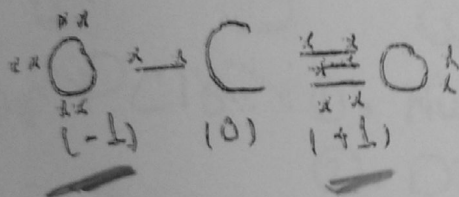


④  $\text{CO}_2$  ∴ Number of valence electrons =  $\text{C}(4) + \text{O}(2 \times 6) = 16e^-$

② Distributing  $16e^-$



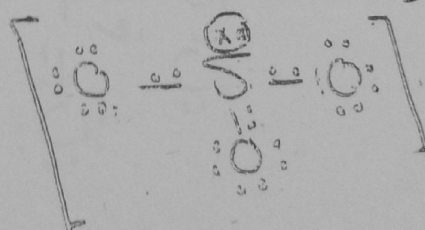
C atom still less than octet so by forming multiple bonds (we need two  $\pi$  bonds) as follows:-



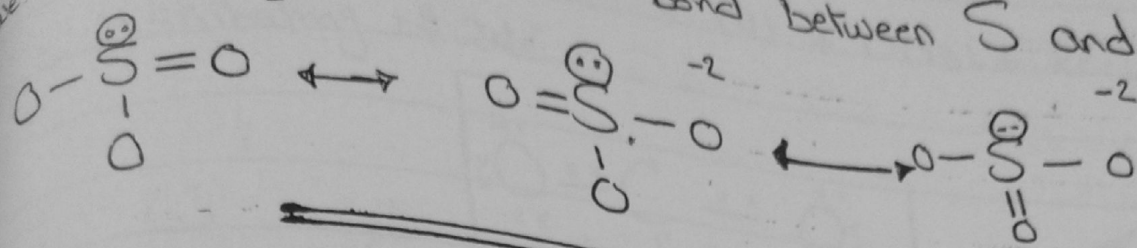
⑤  $\text{SO}_3^{2-}$  ∴ ① Number of valence electrons =  $\text{S}(6) + \text{O}(3 \times 6) +$

$2e^-$  (anion) =  $26e^-$

② Distributing  $26e^-$



According to F.C Calculations to make S and one O atom F.C = Zero we should form one multiple bond between S and O



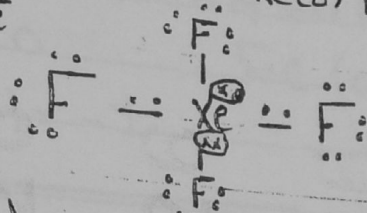
write the Lewis Formula for the following:

① XeF<sub>4</sub> ② SF<sub>4</sub> ③ PCl<sub>5</sub> ④ CO<sub>3</sub><sup>-2</sup> ⑤ NO<sub>3</sub><sup>-</sup> ⑥ FNO<sub>2</sub>

+ ANSWERS

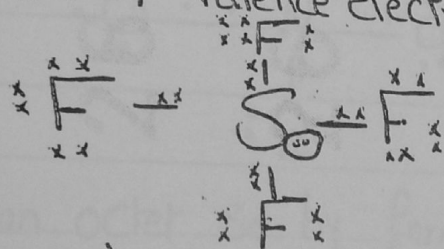
XeF<sub>4</sub> :: ① Number of valence electrons = Xe(8) + F(4x7) = 36e<sup>-</sup>

② Distributing 36e<sup>-</sup>



36e<sup>-</sup> are distributed and two lone pair of e<sup>-</sup> on the Central atom.

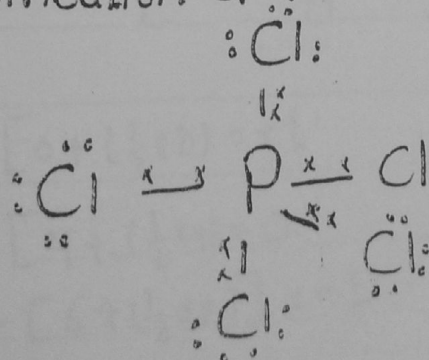
SF<sub>4</sub> :: ① Number of valence electrons = S(6) + F(7x4) = 34e<sup>-</sup>



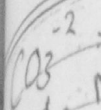
34e<sup>-</sup> are distributed and one lone pair of e<sup>-</sup> on the Central atom.

PCl<sub>5</sub> :: ① + Number of valence e<sup>-</sup> = P(5) + Cl(5x7) = 40e<sup>-</sup>

② + Distribution of 40e<sup>-</sup>

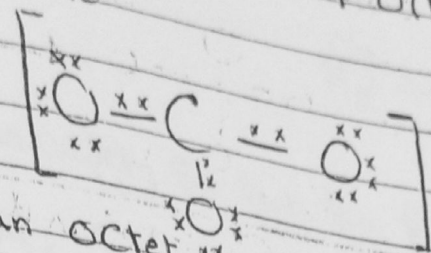




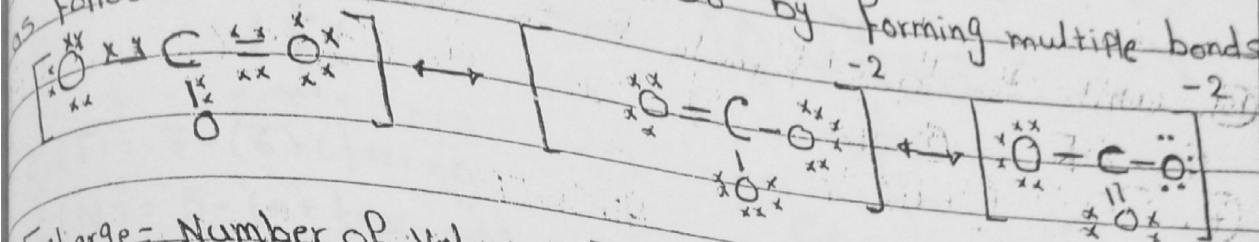


1. Number of valence e =

2. Distributing of 24e  $\text{C}(4) + \text{O}(18) + 2e(\text{anion}) = 24e$



Atom is less than octet. So by forming multiple bonds



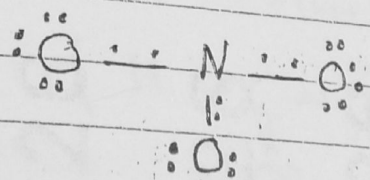
F. charge = Number of valence e

$$\text{F.C. [C]} = 4 - (0 + \frac{1}{2} \times 8) = 0$$

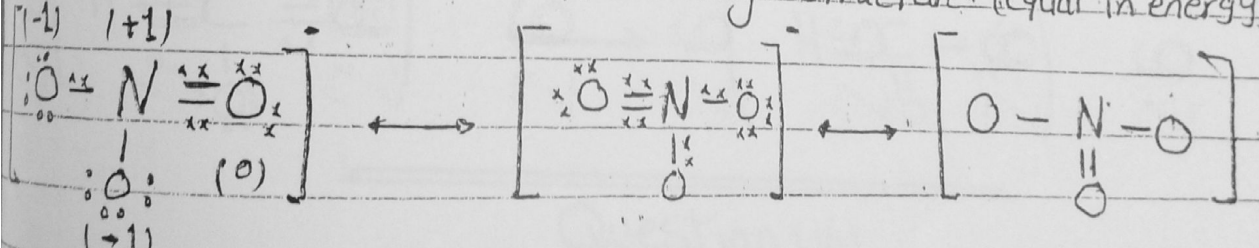
$$\text{F.C. [Single O]} = 6 - (6 + \frac{1}{2} \times 2) = -1$$

$$\text{F.C. (O-double)} = 6 - (4 + \frac{1}{2} \times 4) = 0$$

$\text{NO}_3^-$  : Number of valence e =  $\text{N}(5) + \text{O}(6 \times 3) + 1e(\text{anion}) = 24e$



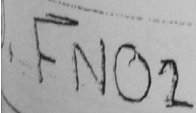
Central atom less than octet so by forming multiple bonds we will have three possible Resonating Structure (equal in energy)



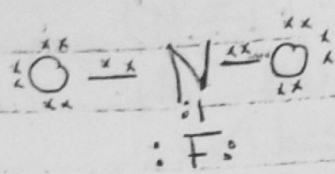
$$\text{F.C. on N} \rightarrow 5 - [0 + (\frac{1}{2} \times 8)] = +1$$

$$\text{O (double)} \rightarrow 6 - [4 + (\frac{1}{2} \times 4)] = 0$$

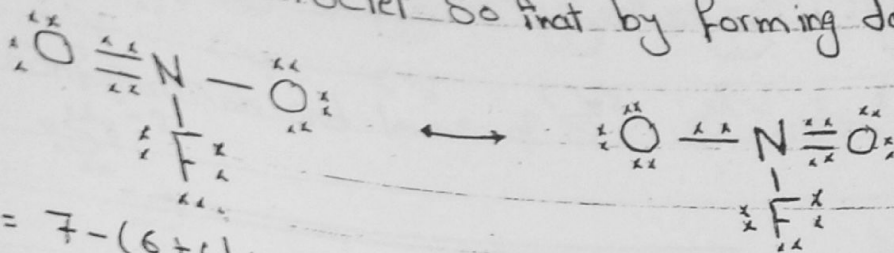
$$\text{O (Single)} \rightarrow 6 - [6 + (\frac{1}{2} \times 2)] = -1$$



- ① Number of Valence  $e^- = F(7) + N(5) + O(6 \times 2) = 24e^-$   
 ② Distributing 24 Valence  $e^-$ :



Central atom is less than octet so that by forming double bond.



$$F.C(F) = 7 - (6 + \frac{1}{2} \times 2) = 0$$

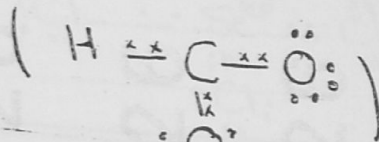
$$F.C(N) = 5 - (0 + \frac{1}{2} \times 8) = +1$$

$$F.C(O) = 6 - (4 + \frac{1}{2} \times 4) = \text{Zero (0)}$$

$$F.C(-O) = 6 - (6 + \frac{1}{2} \times 2) = -1$$

⑦  $\text{SO}_3^{2-}$  (is solved before)

⑧  $\text{HCOO}^-$  : ① Number of Valence  $e^- = H(1) + C(4) + O(6 \times 2) + 1(\text{anion}) = 18e^-$   
 ② Distributing 18e



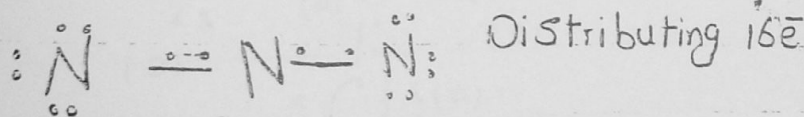
by forming multiple bond because central atom is less than octet.



### Question (16)

① what is the Lewis structure of  $\text{N}_3^-$ ,  $\text{N}_3^{+2}$ ,  $\text{Ca}^{+2}$ ,  $\text{S}^{-2}$

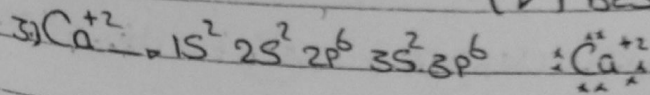
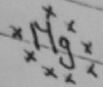
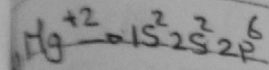
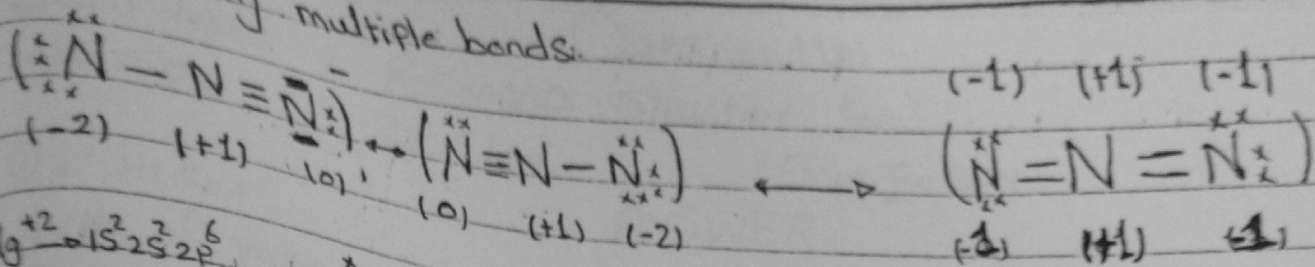
②  $\text{N}_3^- \rightarrow$  Number of Valence  $e^- = N(5 \times 3) + 1(\text{anion}) = 16e^-$



③ N atom still has less than octet.



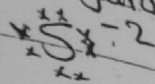
So by forming multiple bonds



✓ best one

Electron Configuration and Lewis of S<sup>-2</sup> S<sup>-2</sup> 1s<sup>2</sup> 2s<sup>2</sup> 2p<sup>6</sup> 3s<sup>2</sup> 3p<sup>6</sup>

Lewis



Draw Lewis Structure For the following

Question (17)

① SeCl<sub>4</sub>

② PSCl<sub>3</sub> (P is Central)

(exam Jan 2012)

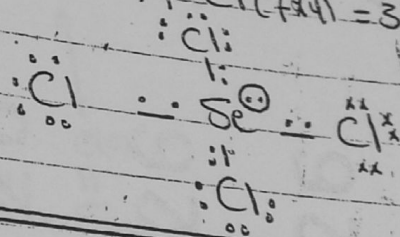
ANSWER

① SeCl<sub>4</sub>

(Se=34)

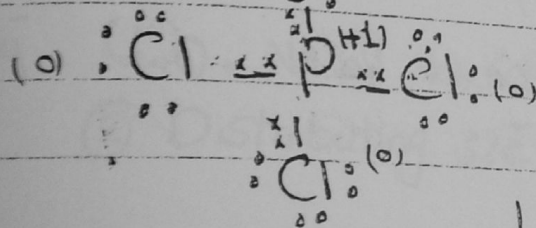
① Number of valence e<sup>-</sup> = Se(6) + Cl(7x4) = 34e<sup>-</sup>

② Distributing 34e<sup>-</sup>



② PSCl<sub>3</sub>

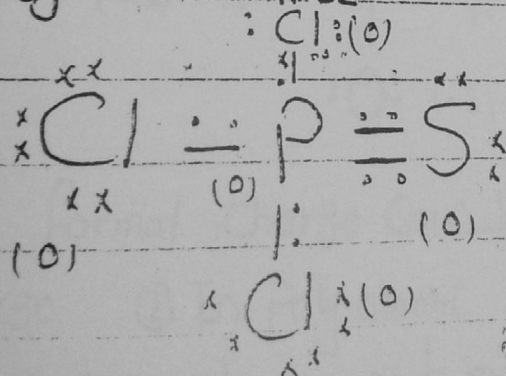
① Number of valence e<sup>-</sup> = P(5) + S(6) + Cl(2x1) = 32e<sup>-</sup>



Making Multiple bond between P and S for Most Suitable

Lewis Structure (lowest energy)

The lowest energy one will be



Question (18)

$$1) \text{H}_2\text{SO}_4$$

2)  $\text{ClO}_4^-$

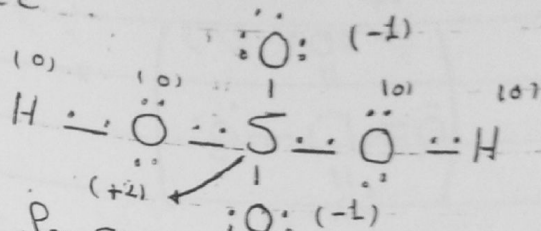
(Exam Jan 2010)

\* ANSWER \*

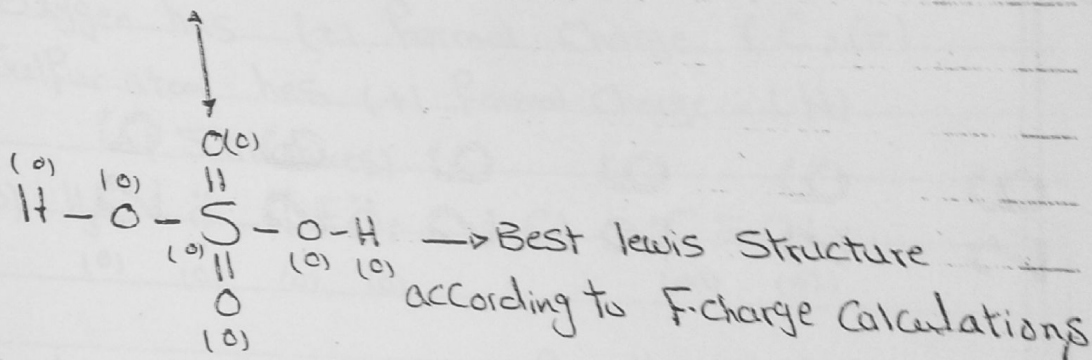
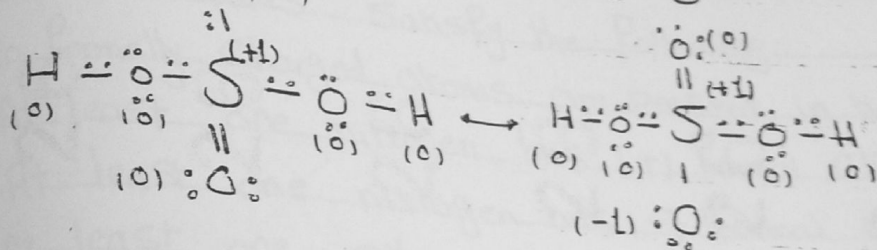
①  $\text{H}_2\text{SO}_4$

① No of Valence

② Distributing 32 e



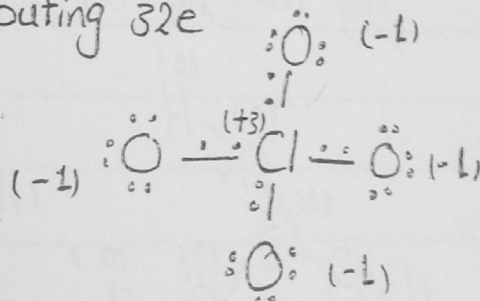
Allowing Resonance for Stabilization and lowering of energy by forming Multiple bonds between S and O



②. ClO<sub>2</sub>

$\therefore \text{①. No of valence } e^- = \text{Cl}(7) + \text{O}(2 \times 6) + 1e^- (\text{anion}) = 32e^-$

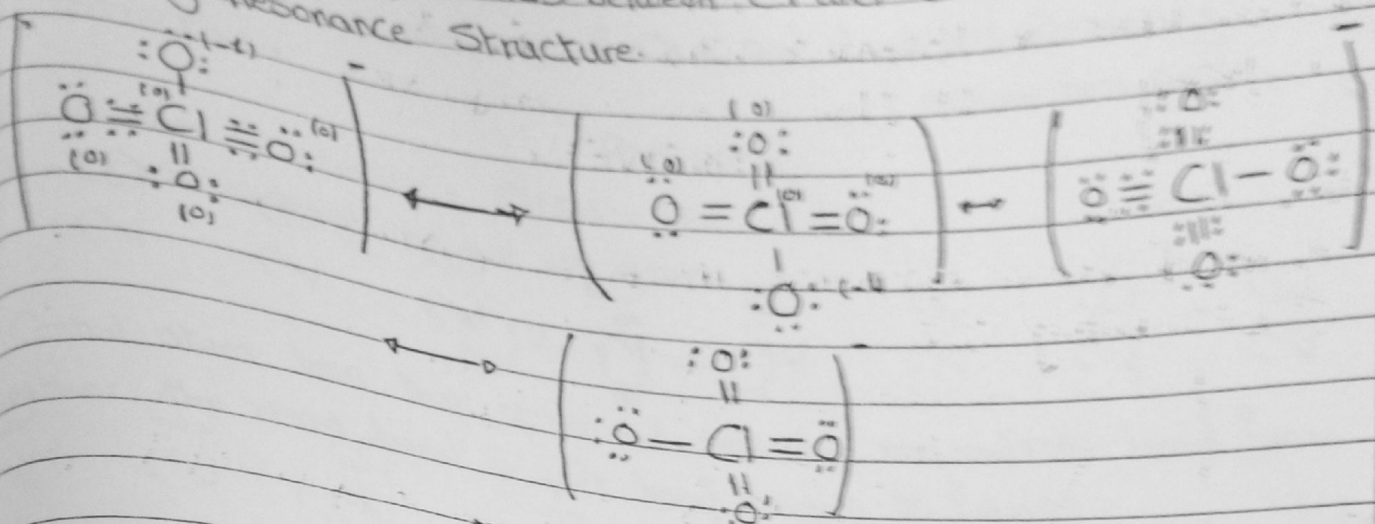
② Distributing  $32e^-$



\* According to Formal Charge Calculations: ① atoms tend to have zero formal charge ② In Polyatomic Ions most electronegative atom acquire formal negative charge and most electropositive atom acquire formal positive charge.

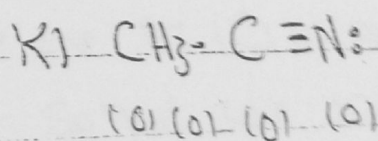
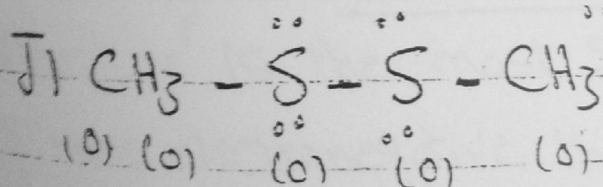
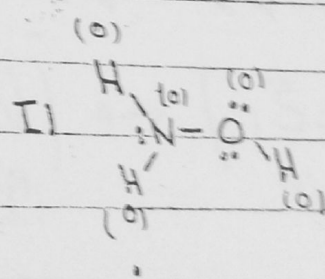
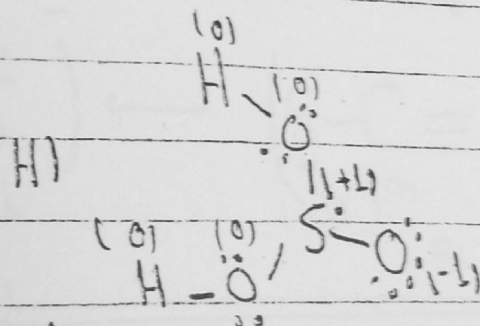
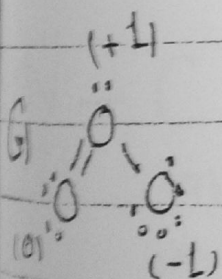
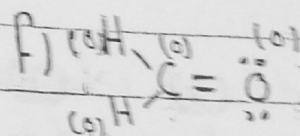
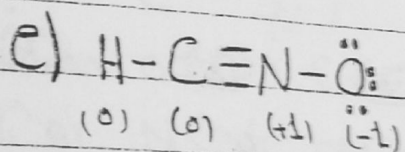
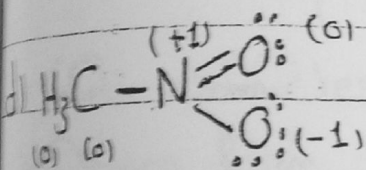
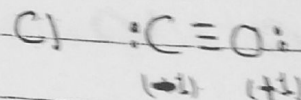
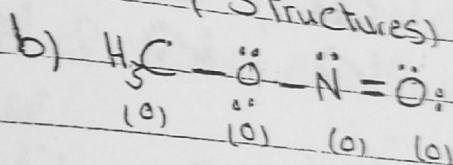
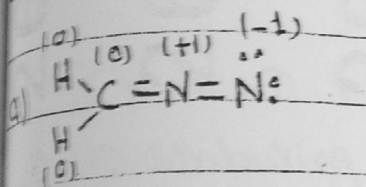


So by forming multiple bonds between Cl and O we obtain the following Resonance Structure.



### Question (19)

- which structures satisfy the following: (Exam June 2012)
- 1) No formally charged atoms are present in the structure (b, f, I, J, K)
  - 2) At least one nitrogen has (+) Formal Charge (a, d, e)
  - 3) At least one oxygen has (-) Formal Charge (a)
  - 4) At least one sulfur atom has (+) Formal Charge (c, G)

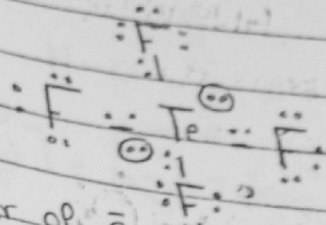


Question (20)



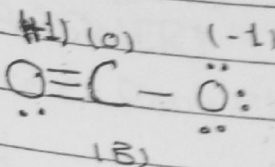
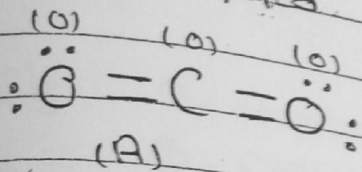
Draw Lewis dot structures for  $\text{TeF}_4^{2-}$  (Te (Z=52), F (Z=9))

No. of Valence  $e^-$ :  $\text{Te}(6) + \text{F}(4 \times 7) + 2(\text{anion}) = 36e^-$



excess two lone pair of  $e^-$  are put on central atom.

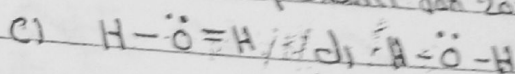
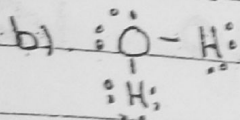
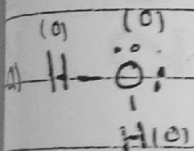
which is preferred? why?



(Exam Jan 2011)

(A) is better because there is no formal charge on any of the atoms.

which of the following has the correct electron-dot formula? why?



(Exam Jan 2011)

(A) is correct because there is no formal charge on any of the atoms.

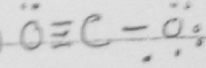
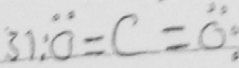
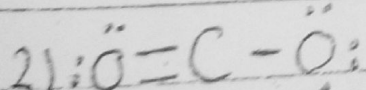
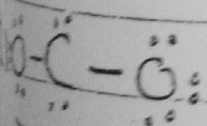
(B) LH has more than octet so x

(C) Oxygen: the more electronegative atom has +F.C so x

(D) Oxygen don't complete its octet.

which of the following is the correct electron-dot formula of  $\text{CO}_2$ ?

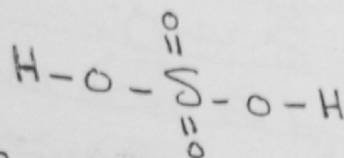
(Exam Jan 2011)



(3) is the correct answer? Say why due to formal charge

Question (23)

ANSWER



$$FC = 6 - [0 + (\frac{1}{2} \times 12)] = 6 - 6 = \text{Zero}$$

Question (24)

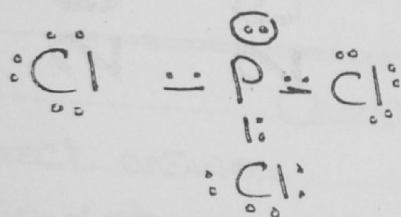
(Jan 2013)

①  $\text{PCl}_3$     ②  $\text{SF}_6$     ③  $\text{SnCl}_4$     ④  $\text{PF}_5$     ⑤  $\text{SO}_2$

[ANSWER]

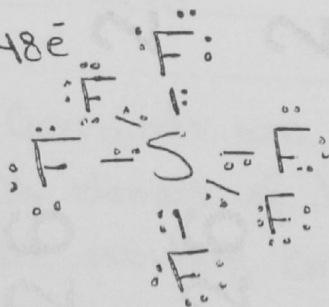
①  $PCl_3$  :-

- ① - Number of valence  $e^- = P(5) + Cl(3 \times 7) = 26e^-$
- ② - Distributing  $26e^-$



(excess of  $\bar{e}$  on Central atom)

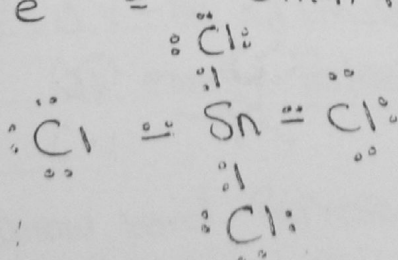
②.  $SF_6$   $\therefore$  Number of valence  $e^- = (S)(6) + F(6 \times 7) = 48e^-$   
 $\therefore$  Distributing  $48e^-$   $\therefore \ddot{F}:$



③.  $\text{SnCl}_4$  (Sn  $1Z=50$ )

Q. Number of valence  $\bar{e}$  =  $5n(4) + Cl(4 \times 7) = 32\bar{e}$

②. Distributing  $32e^-$



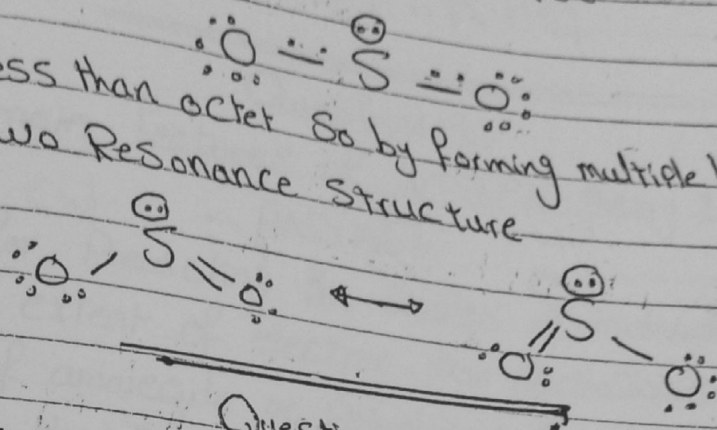
④.  $PF_5$  such as  $PCl_5$  (solved before)



SO<sub>2</sub>

- ①. Number of valence e<sup>-</sup> = S(6) + O(6x2) = 18e<sup>-</sup>
- ②. Distributing 18e<sup>-</sup> =

③. Central atom less than octet so by forming multiple bond from either oxygen we will have two Resonance Structure



Write Formula For the Compounds. Question (25) (Jan 2013)  
Lithium Fluoride, Aluminium oxide

Lithium Fluoride: LiF

Aluminium oxide: Al<sub>2</sub>O<sub>3</sub>

ANSWER

Show how by distributing electrons.

Question (26) (Jan 2013)

Choose the Correct answer:

- ①. The ion that isoelectronic with CO is

1) CN<sup>-</sup> 2) O<sup>+2</sup> 3) O<sup>-2</sup> 4) N<sub>2</sub><sup>+</sup>

- ②. The most stable ionic compounds will result when:

- a) element of low I.E. with element of high electron affinity
- b) when the lattice energy of the resulting compound is very large
- c) both a and b (✓)

③. Electron affinity of O = -141 kJ/mole means:

- 1) energy absorbed 2) energy Released (✓) 3) No Change

NH<sub>3</sub> → BCl<sub>3</sub> the arrow bond Represents:

- 1) ionic 2) Coordinate Covalent (✓) 3) Covalent 4) Covalent and ionic

# FINAL REVISION

## CHAPTER (4)

① Describe the main features of VSEPR theory?

Question (1)

ANSWER

- ①. Theory used for Predicting the shapes of individual molecules Based on their extent of electron-pair electrostatic repulsion.
- ②. The geometry of a molecule or Polyatomic Ion is determined by repulsion among the pairs of electrons associated with central atom.
- ③. These pair of electrons may be bonding or non-bonding (lone pair of e<sup>-</sup>).
- ④. Only valence e<sup>-</sup> of the central atom influence in molecular shape.
- ⑤. In determining molecular shape we should follow this procedure:
  - \* Determine the central atom (lowest subscript, able to form most bonds)
  - \* Draw electron dot structure (Lewis structure) by arranging ligand around central atom.
  - \* Modify geometry to determine the molecular shape (unbonded e<sup>-</sup> will affect the shape of the molecule).

Question (2)

②. According to the VSEPR theory, what are the arrangement of two, three, four, five and six valence-shell electron pairs about an atom?

ANSWER

- \* Draw all possible structures from  $AX_2$  to  $AX_4E_2$  with example & shape at each time. (as shown in table)

Question (3)

③. Why is lone pair expected to occupy an equatorial position instead of an axial position in trigonal bi-pyramidal?



### ANSWER

- Electron Pairs in the equatorial positions form larger angles with their neighbors in the equatorial positions ( $120^\circ$ ), than pairs in the axial positions ( $90^\circ$ ).
- Lone Pairs repel each other more strongly than bonding Pairs so they are found in equatorial positions where their Repulsive force are reduced.

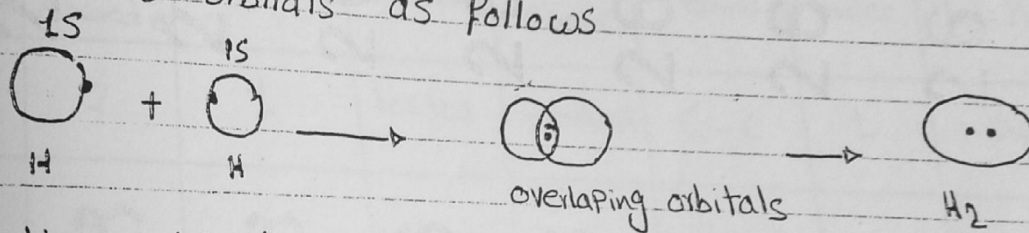
### Question (4)

- Describe the formation of Covalent bond in  $H_2$  from atoms? what does it mean to say that the bonding electrons are shared by the two atoms?

### ANSWER

- Covalent bond is a chemical bond result from overlap of one or an atomic orbital of one atom with atomic orbital of another atom (Sharing of  $e^-$ ).

- In  $H_2$  Molecule: is formed from two H atoms each H has a single electron in  $1s$  orbital. So H-H bond result from overlap of the two  $1s$ -orbitals as follows



H in  $H_2$   $1s \uparrow \downarrow$

- We mean by  $e^-$  are shared by the atoms: Pair of electrons that form covalent bond is shared between the two atoms in the region of overlap as formation of bond means that two orbitals share some common region in space so two electrons or electron density is concentrated in overlap region.

### Question (5)

5. What is the relationship between bond order & bond length? Use an example to illustrate it?

### ANSWER

1. Bond order: determines number of bonds between pairs of atoms in a molecule (No. of bonds that any atom can form is determined by bond order) by counting electron pair in bonding & anti-bonding molecular orbitals

For ex: Bond order of C in  $C-C=1$   
 $C=C=2$

2. Covalent bond can be single, double and triple

- Single bond with bond order = 1 ( $\sigma$  bond) is the longest.

- Triple bond with bond order = 3 ( $1\sigma+2\pi$ ) is the shortest

• Bond length is the internuclear separation of the two bonded atoms so as number of electrons shared between atoms (bond order) increases, bond length decreases (bond strength increases).

→ illustrating with examples:

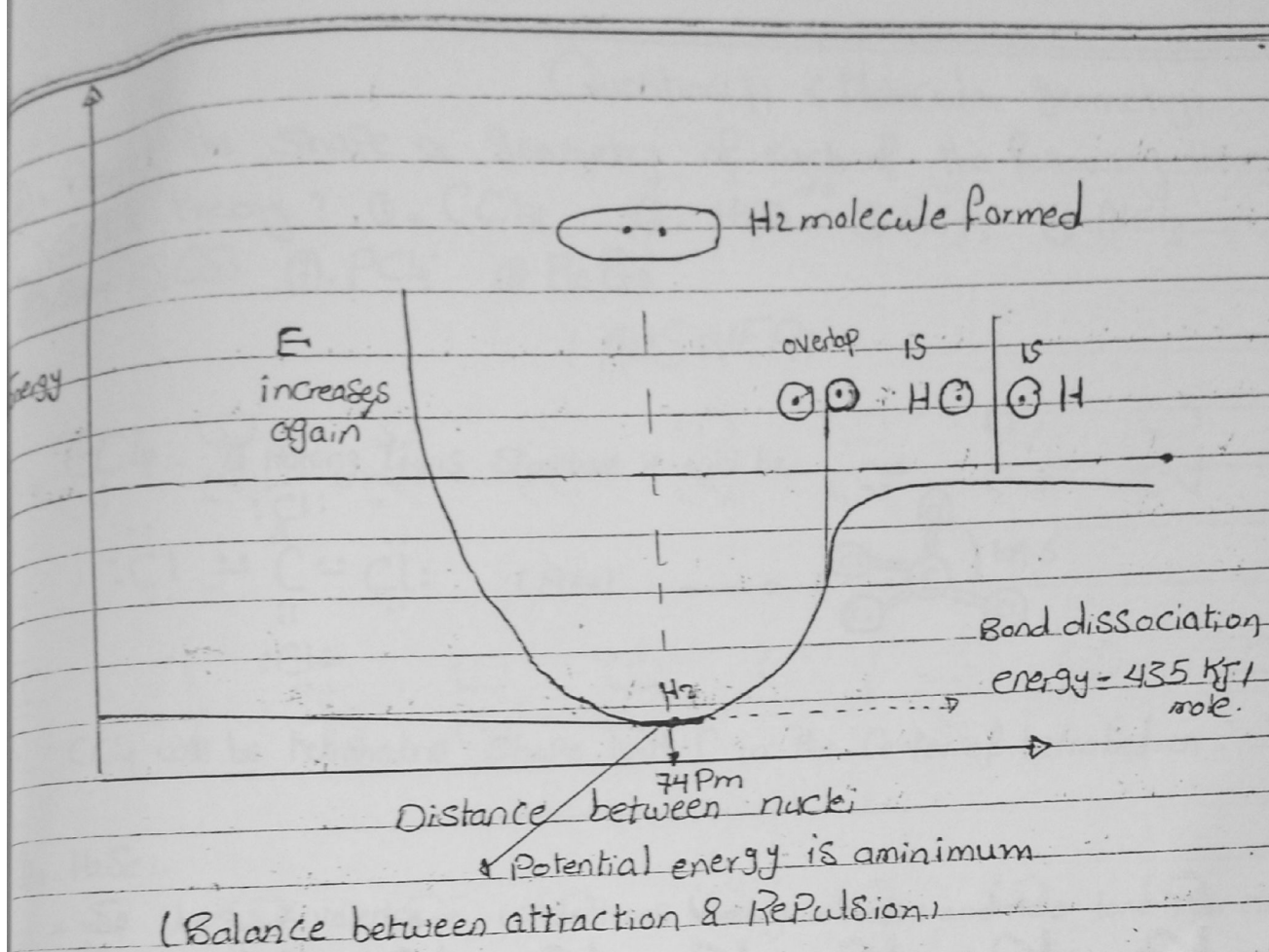
Bond	no of shared e	bond order	bond length	bond strength	Bond	order	length	Strength
Single	2	1	longest	Weakest	$C-C$	1	1.54 Å	348 kJ/mole
double	4	2	medium length		$C=C$	2	1.34 Å	614 kJ/mole
triple	6	3	Shortest	Strongest	$C\equiv C$	3	1.20 Å	839 kJ/mole

### Question (6)

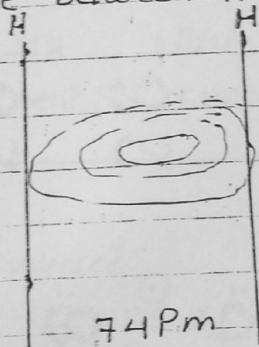
6. Draw potential energy diagram for a molecule such as  $H_2$ , on the diagram indicate bond length & bond strength?

(ANSWER)





\* **Bond length:** The distance between nuclei of the atom when the energy is minimum



\* **Bond energy:** The energy must be supplied to separate atoms from molecules.

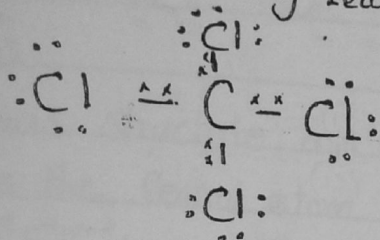
Bond H-H energy =  $435 \text{ KJ/mole}$

\* The two electrons are shared by both atoms and serve as a sort of glue cementing the atoms together.

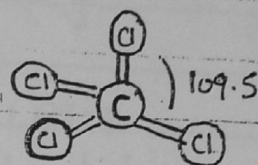
Question (7) < Molecular Geometry >  
 Predict the shape or geometry of each of the following using VSEPR theory? ①  $\text{CCl}_4$  ②  $\text{H}_2\text{Se}$  ③  $\text{AsF}_3$  ④  $\text{AlCl}_3$   
 ⑤  $\text{SiH}_4$  ⑥  $\text{CS}_2$  ⑦  $\text{PCl}_5$  ⑧  $\text{BeBr}_2$

(ANSWER)

①  $\text{CCl}_4$  : Making Lewis Structure it will be:



(AX<sub>4</sub>)

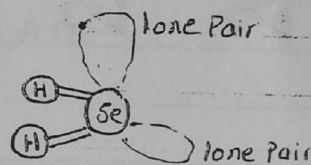
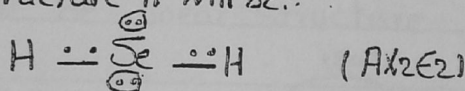


$\text{CCl}_4$  will be tetrahedral shape with C in the center of tetrahedron

②  $\text{H}_2\text{Se}$  :

Se has 6 valence e<sup>-</sup> (2 electrons bonding with H) and two lone pair electrons

• Making Lewis Structure it will be:

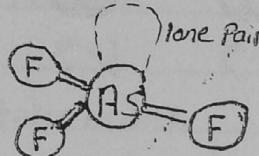


•  $\text{H}_2\text{Se}$  will be nonlinear (bent structure) or V shaped

③  $\text{AsF}_3$  :

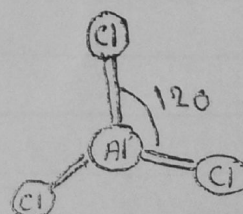
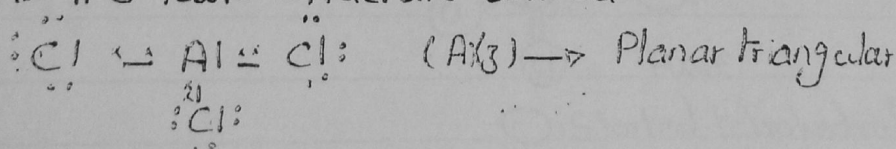
As: [Ar],  $4s^2, 3d^{10}, 4p^3 \rightarrow 5$  valence e<sup>-</sup>

• Lewis Structure will be:  $\text{:}\ddot{\text{F}}\text{---}\text{As}\text{---}\ddot{\text{F}}\text{:}$  (AX<sub>3</sub>E)



•  $\text{AsF}_3$  : Pyramidal Shape

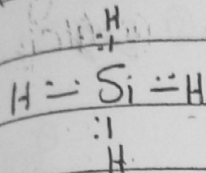
④  $\text{AlCl}_3$  : its Lewis Structure will be:





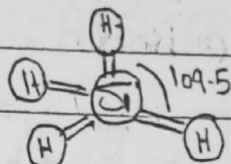
SiH<sub>4</sub>

Si (4 valence e<sup>-</sup>) So Lewis structure will be:



AX<sub>4</sub>

it will be tetrahedral shape

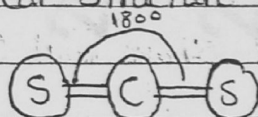


CS<sub>2</sub>

Lewis structure will be:  $\text{S} \equiv \text{C} \equiv \text{S}$  (AX<sub>2</sub>)

C is the central atom with 5 atom doubly bonded to it

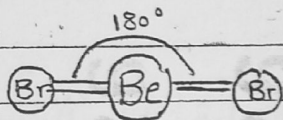
CS<sub>2</sub> (AX<sub>2</sub>) will have linear structure.



BeBr<sub>2</sub>

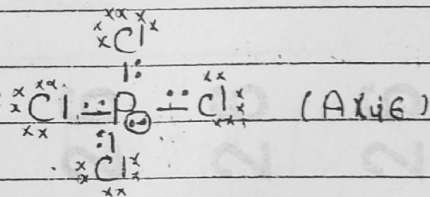
Lewis structure will be:  $\text{Br} - \text{Be} - \text{Br}$  (AX<sub>2</sub>)

BeBr<sub>2</sub> (AX<sub>2</sub>) will be linear structure.

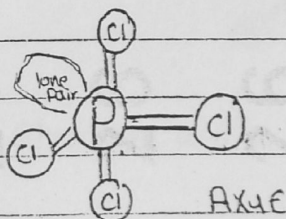


③ PCl<sub>4</sub><sup>-</sup>

Lewis structure will be:



(AX<sub>4</sub>E) So its shape will be distorted tetrahedral.



Distorted tetrahedral

### Question (8)

③. Predict the geometry of each of the following, using VSEPR theory?

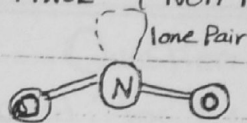
- ①.  $\text{NO}_2^-$     ②.  $\text{NH}_2^-$     ③.  $\text{ClO}_2^-$     ④.  $\text{H}_3\text{O}^+$     ⑤.  $\text{BeF}_3^-$     ⑥.  $\text{BrO}_3^-$

• ANSWER •

①.  $\text{NO}_2^-$

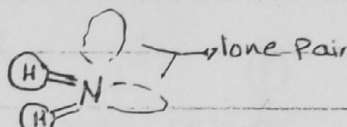
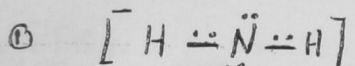
• Lewis structure will be  $\left[ \begin{array}{c} \cdot\ddot{\text{O}}::\ddot{\text{N}}::\ddot{\text{O}}\cdot \\ \cdot\cdot\cdot\cdot \end{array} \right]^- \longleftrightarrow \left[ \begin{array}{c} \cdot\ddot{\text{O}}::\ddot{\text{N}}::\ddot{\text{O}}\cdot \\ \cdot\cdot\cdot\cdot \end{array} \right]^-$

• It will be of the type  $\text{AX}_2\text{E}$  (Non linear) or bent or V-Shaped



②.  $\text{NH}_2^-$

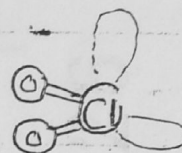
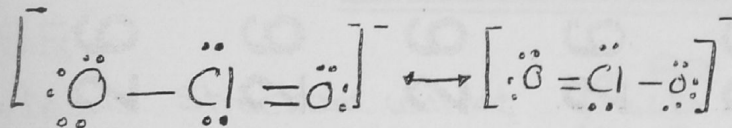
• Lewis structure will be:



②. It is of the type  $\text{AX}_2\text{E}_2$  (nonlinear) or bent or V-Shaped structure

③.  $\text{ClO}_2^-$

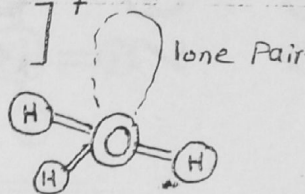
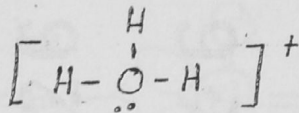
• Lewis structure will be:



• It will be of the type  $\text{AX}_2\text{E}_2$  So it will be linear or bent or V-Shaped.

④.  $\text{H}_3\text{O}^+$

• Lewis structure will be

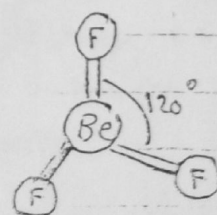


• It is of the type  $\text{AX}_3\text{E}$

• It has a pyramidal shape

⑤.  $\text{BeF}_3^-$

→ its Lewis structure will be  $\left( \text{F}-\underset{\text{F}}{\text{Be}}-\text{F} \right)^-$



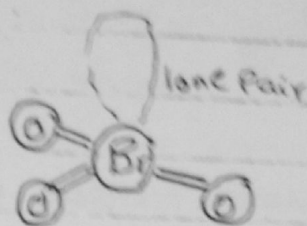
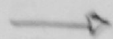
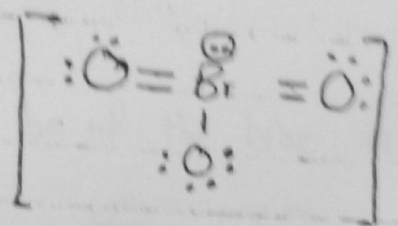
• It is of the type  $\text{AX}_3$  So it will be planar triangular

nonlinear



Q.  $\text{BrO}_3^-$

According to Lewis structure :-



It is of the type  $\text{AX}_3\text{E}$  So it has pyramidal shape

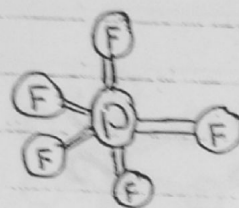
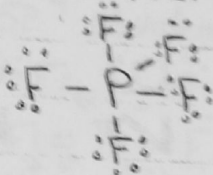
Question 19)

Q. what geometry is expected for the following according to VSEPR ?  
 ①.  $\text{PF}_5$  ②  $\text{BrF}_3$  ③  $\text{BrF}_5$  ④  $\text{SCl}_4$  ⑤  $\text{ClF}_3$  ⑥  $\text{SeF}_4$  ⑦  $\text{TeF}_6$   
 ⑧  $\text{SnCl}_3$  ⑨  $\text{PCl}_5$  ⑩  $\text{GaF}_3$  ⑪  $\text{PH}_3$

★ ANSWER ★

①.  $\text{PF}_5$  :-

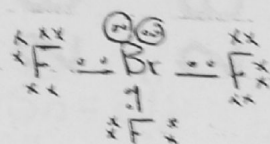
• Lewis structure will be



It is of the type  $\text{AX}_5$  So it will be trigonal bipyramidal.

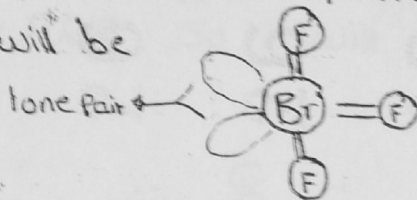
②.  $\text{BrF}_3$  :-

• Lewis structure will be



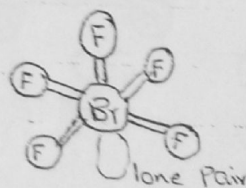
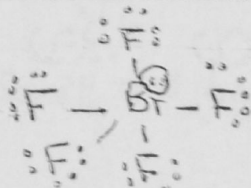
It is of the type  $\text{AX}_3\text{E}_2$  So it is T-shaped.

• Molecular geometry will be



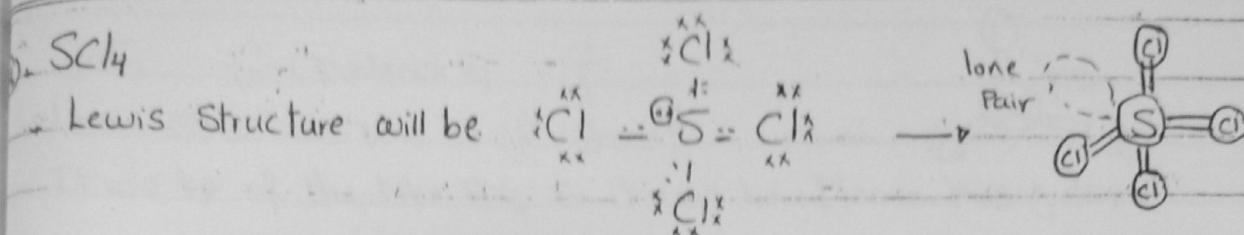
③.  $\text{BrF}_5$  :-

• Lewis structure will be



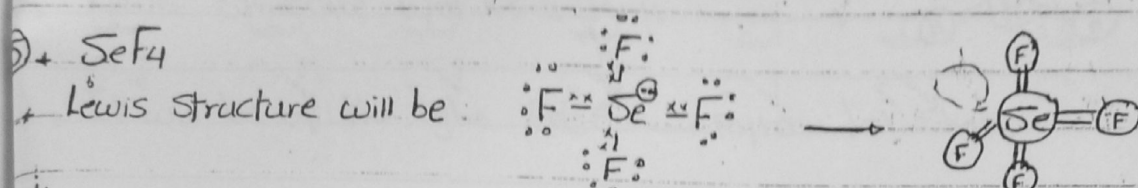
It is of the type  $\text{AX}_5\text{E}$  So it has square pyramidal shape

⑤  $\text{SCl}_4$



It will be of the type  $\text{AX}_4\text{E}$  so it will be distorted tetrahedral

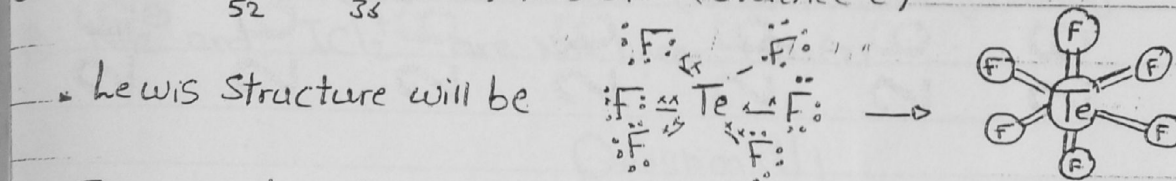
⑥  $\text{SeF}_4$



It will be of the type  $\text{AX}_4\text{E}$  so it will be distorted tetrahedral

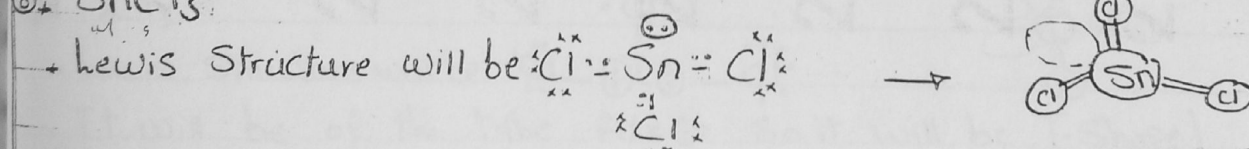
⑦  $\text{ClF}_5$  will also have the same shape as  $\text{BrF}_5$  (square pyramidal)

⑧  $\text{TeF}_6$   $\text{Te} \rightarrow \text{Kr}, 5s^2, 4d^{10}, 5p^4$  (6 valence e)



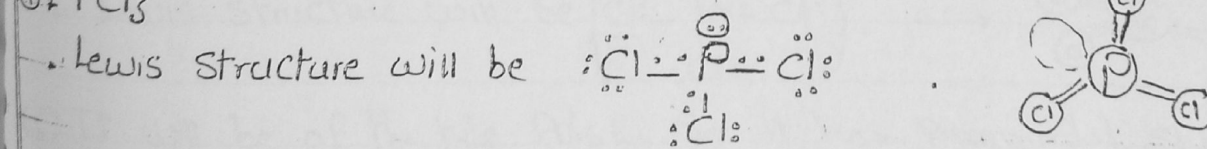
It will be of the type  $\text{AX}_6$  so it will be octahedral shape

⑨  $\text{SnCl}_3$



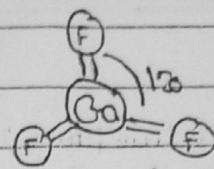
It will be of the type  $\text{AX}_3\text{E}$  so it will be pyramidal shape

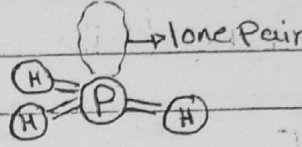
⑩  $\text{PCl}_3$



It will be of the type  $\text{AX}_3\text{E}$  so it will be pyramidal shape



$\text{GaF}_3$  Ga (3 valence e)  $\text{F} \cdot \cdot \text{F} \cdot$   
 Lewis structure will be  $\text{F} \cdot \cdot \text{Ga} \cdot \cdot \text{F} \cdot$   $\rightarrow$    
 It will be of the type  $\text{AX}_3$  so it will be Planar triangular

$\text{PH}_3$   
 Lewis structure will be  $\text{H} \cdot \cdot \text{P} \cdot \cdot \text{H}$   $\rightarrow$    
 It will be of the type  $\text{AX}_3\text{E}$   $\rightarrow$  Pyramidal shape

### Question (10)

which of the following molecules or ions is linear?  
 ①  $\text{BeCl}_2$  ②  $\text{NH}_2^-$  ③  $\text{CS}_2$  ④  $\text{ICl}_2^+$

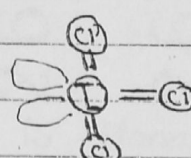
### ANSWER

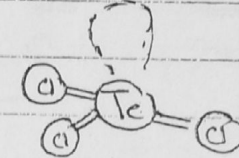
①  $\text{BeCl}_2$  and  $\text{CS}_2$  are linear (draw Molecular geometry)  
 ②  $\text{NH}_2^-$  and  $\text{ICl}_2^+$  are Non linear ( $\text{AX}_2\text{E}_2$ )

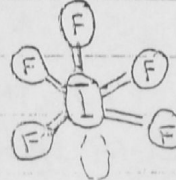
### Question (11)

what molecular geometry would you expect for  $\text{ICl}_3$ ,  $\text{TeCl}_3$  and  $\text{IF}_5$

### ANSWER

①  $\text{ICl}_3$   
 Lewis structure will be  $\cdot \ddot{\text{Cl}} \cdot - \text{I} - \cdot \ddot{\text{Cl}} \cdot$   $\rightarrow$    
 It will be of the type  $\text{AX}_3\text{E}_2$  so it will be T-Shaped

②  $\text{TeCl}_3^+$   
 Lewis structure will be  $(\cdot \ddot{\text{Cl}} \cdot - \text{Te} - \cdot \ddot{\text{Cl}} \cdot)^+$   $\rightarrow$    
 It will be of the type  $\text{AX}_3\text{E}$  so it has pyramidal shape

③  $\text{IF}_5$   
 Lewis structure will be  $\text{F} - \text{I} - \text{F}$   $\rightarrow$    
 it will be of the type  $\text{AX}_5\text{E}$   
 so it will be Square pyramidal.

### Question (12)

2. Choose the correct answer:

- ①. Carbon tetrachloride ( $\text{CCl}_4$ ) has no net dipole moment because of  
 (A) It's planar structure (B) it's regular tetrahedral structure (✓)  
 (C) Similar sizes of carbon and chlorine atoms  
 (d) Similar electron affinity of C and Cl

②. The types of bonds present in  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  are only

- (A) electrovalent and covalent (B) electrovalent and coordinate covalent  
 (C) electrovalent, covalent and coordinate covalent (✓)  
 (d) covalent and coordinate covalent

③. The hydrogen bond is strongest in

- (1)  $\text{O}-\text{H} \cdots \text{S}$  (2)  $\text{S}-\text{H} \cdots \text{O}$  (3)  $\text{F}-\text{H} \cdots \text{O}$  (4)  $\text{F}-\text{H} \cdots \text{S}$  (✓)

④. The hybridization of sulfur dioxide is  $\text{AX}_2\text{E}$

- (1)  $\text{sp}$  (2)  $\text{sp}^2$  (3)  $\text{sp}^3$  (4)  $\text{dsp}^2$

### Question (13)

③. Put true or false:

①. The shape of  $\text{CH}_3^+$  is tetrahedral (X)  $\Rightarrow \text{CH}_3^+$  is planar triangular

②. All molecules with polar bonds have dipole moment (X)

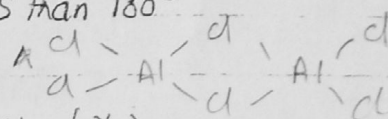
Reason: if symmetrical arrangement of atoms is present, there is no dipole moment for example  $\text{CCl}_4$ .

③.  $\text{SnCl}_2$  is non linear molecule (✓)

There is lone pair around central atom. Due to repulsion between lone pair and bonding pair the bond angle is less than  $180^\circ$

④. All the Al-Cl bonds in  $\text{Al}_2\text{Cl}_6$  are equivalent (X)

terminal Al-Cl bonds are different from bridge Al-Cl bonds.





③. Dipole moment of  $\text{CH}_3\text{F}$  is greater than  $\text{CH}_3\text{Cl}$ . (X)

Bond distance is larger in  $\text{CH}_3\text{Cl}$  ( $\text{Cl} > \text{F}$  in size) and dipole moment  
= Charge  $\times$  distance

④. Nitric oxide, though an odd electron molecule is diamagnetic in liquid state (X)

In liquid state, some dimerization takes place but it is only partial so it is still paramagnetic only.

⑤. The presence of polar bonds in polyatomic molecules suggest that molecule has non-zero dipole moment (X)

→ Some polyatomic molecules having polar bonds have zero dipole moment

⑥.  $\text{H}_2\text{O}$  is linear molecule (X) is V-shaped or non-linear

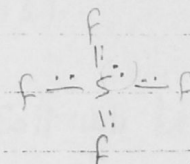
⑦. The basic nature of hydroxides of group 13 (III B) decreases progressively down the group (X)

Basic nature increases as the element become more metallic  
As one goes down a group metallic character increases

⑧. Both potassium ferrocyanide and potassium ferricyanide are diamagnetic (X)

Potassium ferrocyanide is diamagnetic (no unpaired  $e^-$ ) while potassium ferricyanide is paramagnetic (one unpaired  $e^-$ )

⑨. What type of hybridization of S in  $\text{SF}_4$ ?



AX<sub>4</sub>E

→  $\text{sp}^2$  (X)  $\text{sp}^3\text{d}$  (✓)  $\text{sp}^3\text{d}^2$  (X)  $\text{sp}^3$  (X)

⑩. What is the type of hybridization in  $\text{XeF}_4$ ?

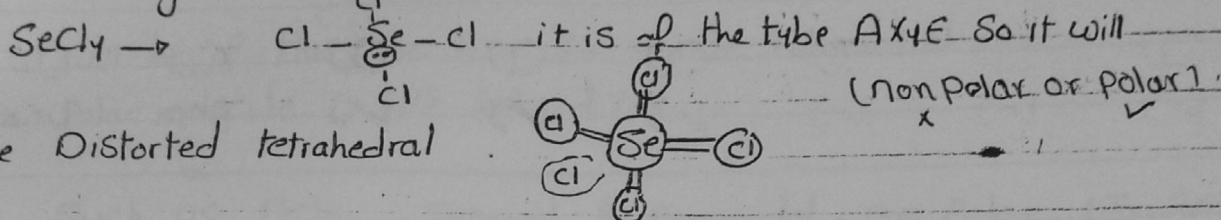
①  $\text{sp}^3$  ②  $\text{sp}^2$  ③  $\text{sp}^3\text{d}^2$  (✓) ④  $\text{sp}^3\text{d}$

### Question (14)

14. Sketch the shape and determine the type of hybridization in each of the following: ①.  $\text{SeCl}_4$  ②.  $\text{PSCl}_3$  (exam June 2013)

### ANSWER

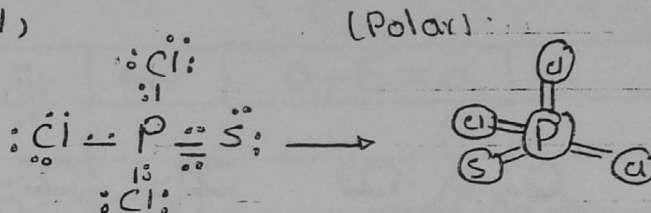
According to Lewis Structure



$\text{AX}_4\text{E}$  belongs to  $\text{AX}_5$  so  $\text{sp}^3\text{d}$  hybridized.

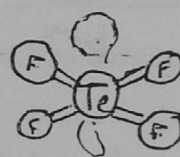
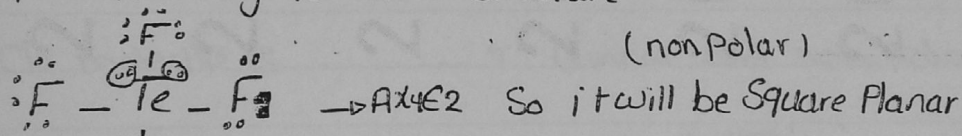
2.  $\text{PSCl}_3$  : (P is Central)

Lewis Structure will be



it is of the type  $\text{AX}_4$  So it has a tetrahedral shape  
 $\text{AX}_4$  So  $\text{sp}^3$  hybridized

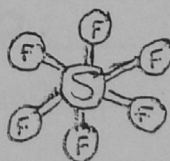
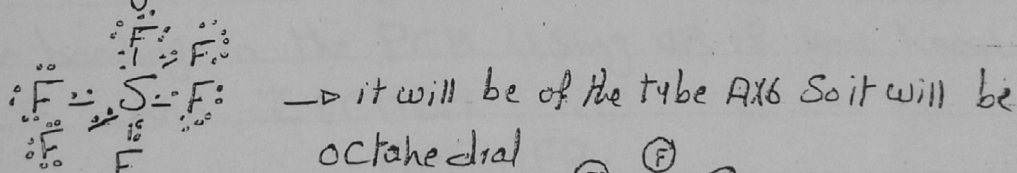
③.  $\text{TeF}_4^{-2}$  : According to Lewis Structure



it will be  $\text{sp}^3\text{d}^2$  hybridized.

④.  $\text{SF}_6$  : According to Lewis Structure

Jan (2013)



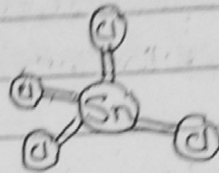
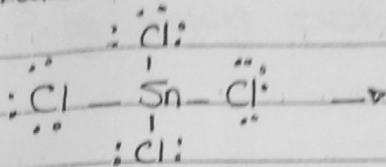
$\text{sp}^3$  hybridized (non polar)



$\text{SnCl}_4$  (4 valence  $e^-$  for Sn)

(Jan 2013)

Lewis structure will be:

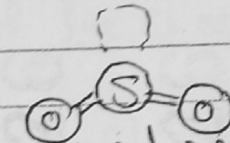
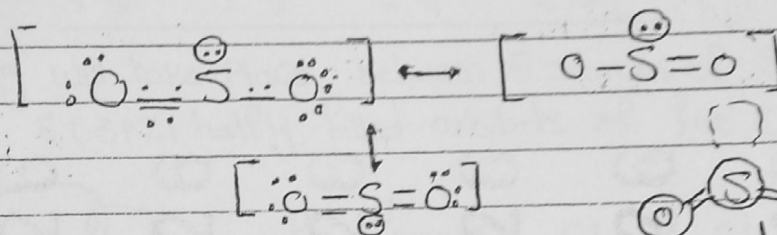


it will be of the type  $\text{AX}_4$ . So will be tetrahedral non polar molecule,  $\text{sp}^3$  hybridized

PF<sub>5</sub> Such as  $\text{PCl}_5$  → Trigonal bi Pyramidal, non polar (Jan 2013)

$\text{SO}_2$  Lewis structure will be:

(Jan 2013)



V-shaped & nonlinear or bent structure

it will be of the type  $\text{AX}_2\text{E}$

polar molecule with  $\text{sp}^2$  hybridized.

⑧  $\text{BeCl}_2$  molecule is \_\_\_\_\_ (Linear and non polar) (Jan 2013)

⑨ The H-N-H bond angle in  $\text{NH}_3$  is \_\_\_\_\_ (shorter than) The H-O-H bond angle in  $\text{H}_2\text{O}$  (Jan 2013)

Question 115

(Jan 2012)

⑤ Explain the bonding in the  $\text{PCl}_5$  using VB if you know that  $Z(\text{P}) = 15$  and  $Z(\text{Cl}) = 17$

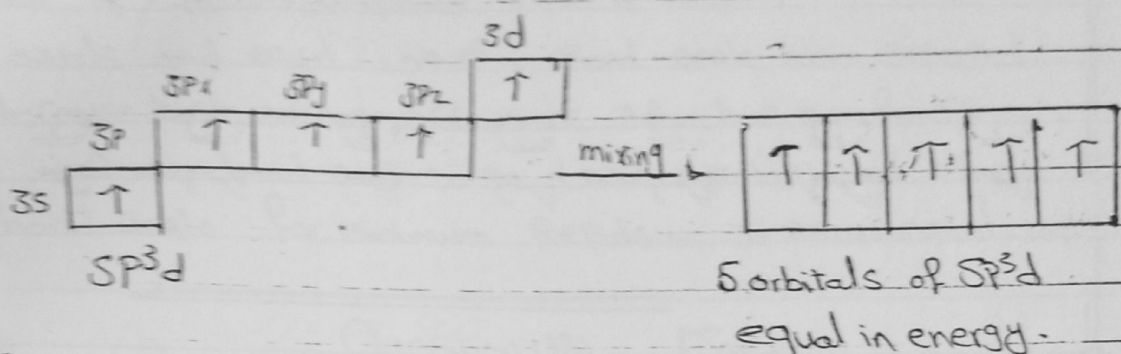
ANSWER

P →  $1s^2 2s^2 2p^6 3s^2 3p^3$  has five valence  $e^-$

$3p_x$	$3p_y$	$3p_z$
↑	↑	↑

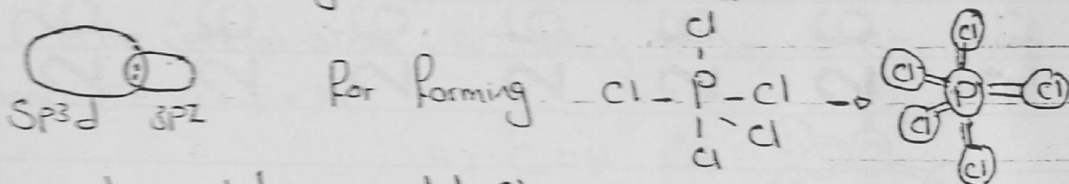
3s ↑↓

So for making five bonds with Cl in  $\text{PCl}_5$  and as Covalent bond is formed between half-filled orbitals, we need five half-filled orbitals in P by making excitation followed by hybridization.



P can form hypervalent compounds due to ability of filling 3d subshell.

Overlapping will take place between 5 orbitals of  $\text{sp}^3\text{d}$  type of P atom with 5 ( $\text{sp}^3$ ) half-filled orbitals of five Cl atoms.

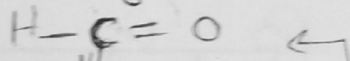


which has trigonal bipyramidal shape

### Question (16)

(June 2011)

Put true or false  
 Q.  $\pi$  bond forms due to the overlap of hybrid orbitals? Give an example  
 (X) due to side by side overlapping between p orbitals as in  $\text{H}_2\text{C}=\text{CH}_2$



Q. Hybridization on oxygen in  $\text{CH}_2\text{O}$  is  $\text{sp}^3$  but on Silicon in  $\text{SiO}_4^{4-}$  is  $\text{sp}^3$  (X)

on oxygen is  $\text{sp}^2$  but on Si in  $\text{SiO}_4^{4-}$  is  $\text{sp}^3$

Q. P often forms compounds  $\text{P}_5$  what hybrid orbitals does P use?  $\text{sp}^3\text{d}$  (explain as in question No 15) (June 2011)



### Question (17) (June 2011)

What do you expect the bond angles to be in  $\text{NH}_4^+$  ion?  
 All will be the same as fourth bond is coordinate covalent and formed of  $2e^-$  transferred from N to  $\text{H}^+$

All are equal and equal  $109.5^\circ$ . Bond angle also expected to be  $90^\circ$  because best overlap occurs at  $90^\circ$ . but two forces work one decrease bond angle to  $90^\circ$  (best overlap) and other increase bond angle for minimum repulsion. So All are equal  $= 109.5^\circ$

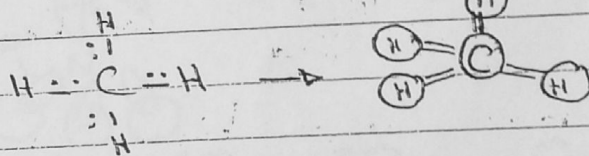
### Question (18) (Jan 2011)

What geometry and hybridization are expected for  $\text{CH}_4$ ,  $\text{NH}_3$ , and  $\text{H}_2\text{O}$  molecules using VSEPR and VB?

### (ANSWER)

①  $\text{CH}_4$ : C is the central atom, according to Lewis structure

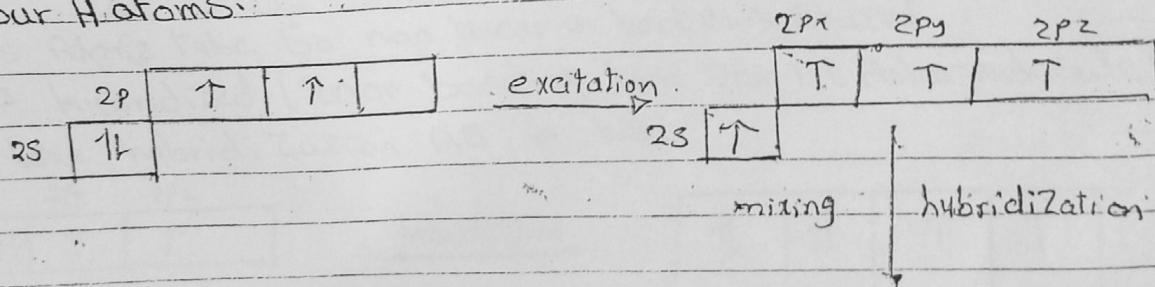
No. of valence  $e^- = 4 + 4 = 8e^-$



No lone pair of  $e^-$  on central atom so it is of the type  $\text{AX}_4$  and will be tetrahedral geometry with bond angle  $109.5^\circ$  (non polar)

C atom in  $\text{CH}_4$  is  $\text{sp}^3$  hybridized

C atom needs four half-filled atomic orbitals for overlapping with four H atoms.



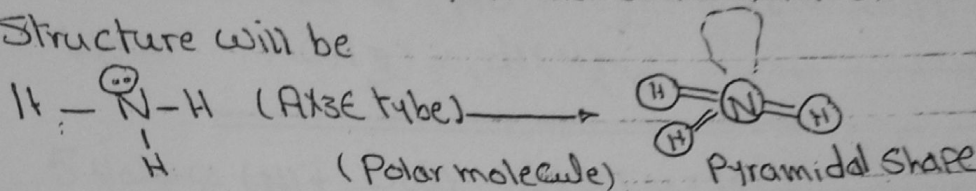
4 H's of H  
 ○ ○ ○ ○

overlap  $4\text{sp}^3$   $\uparrow$   $\uparrow$   $\uparrow$   $\uparrow$

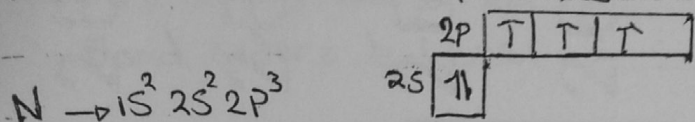
②  $\text{NH}_3$

→ N is the Central atom • No of Valence  $\bar{e} = \text{N}(5) + \text{H}(3) = 8\bar{e}$

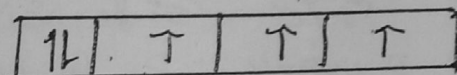
→ Lewis Structure will be



→  $\text{NH}_3$  is also  $\text{sp}^3$  hybridized

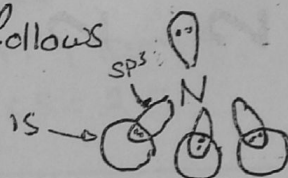


three halfly filled orbitals for making 3 bonds with H but since hybridized orbitals are stronger in overlapping so by mixing (2s, 2p) orbitals we obtain 4 hybridized  $\text{sp}^3$  orbitals as follows:



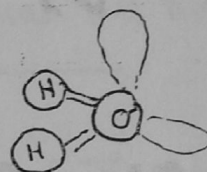
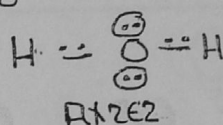
3  $\text{sp}^3$  is halfly filled and ready for overlap while one is completely filled.

• Bonding will occur as follows



③  $\text{H}_2\text{O}$  : according to Lewis Structure

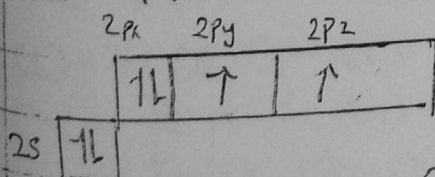
$\bar{e} = \text{O}(6) + \text{H}(2) = 8$



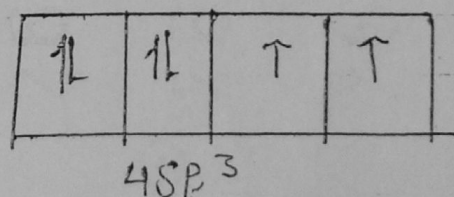
→ it is AXE2 type So non linear or bent or V-Shaped

→  $\text{sp}^3$  hybridized (2 non bonding lone pair) (Polar molecule)

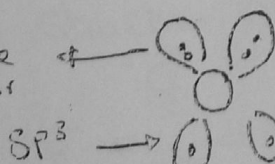
→ Make hybridization AS in  $\text{NH}_3$



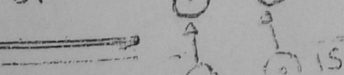
hybridization  $\rightarrow$



2 lone pair



$\text{sp}^3$



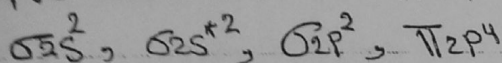


### Question (19)

① what is the bond order of  $\text{NO}^+$ ? (Jan 2011)

- a) (1.5)    b) (2)    c) (3)    d) (4)

$$\text{NO}^+ \bar{e} = 5 \text{ Valence (N)} + 6(1) - 1 (\text{cation}) = 10\bar{e}$$



$$\text{Bond order} = \frac{1}{2} (8 - 2) = 3$$

② what is the molecular geometry of  $\text{NH}_3$ ?

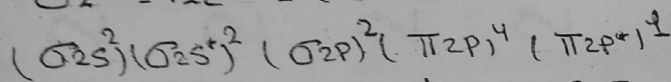
- a) trigonal planar    2) trigonal planar    3) pyramidal (✓)

③ what is the hybridization of S in  $\text{SF}_4$ ? ( $\text{sp}^3\text{d}$ )

④ what is Formal Charge of S in  $\text{H}_2\text{SO}_4$ ? (Zero)

⑤ what is the bond order in  $\text{O}_2^+$ ? (2.5)

$$\text{O}_2 = 12e - 1 (\text{cation}) = 11$$



$$\text{B.O} = \frac{1}{2} (8 - 3) = \frac{1}{2} \times 5 = 2.5$$

⑥ what is the electron pair geometry of central oxygen of ozone  $\text{O}_3$ ?

- 1) linear    2) trigonal planar    3) tetrahedral    4) trigonal bipyramidal

### Question (20)

① Sketch the shape of the following: (Jan 2012)

- ①.  $\text{XeOF}_2$     ②.  $\text{ClOF}_3$     ③.  $\text{SOCl}_2$     ④.  $\text{S}_2\text{O}_4^{2-}$

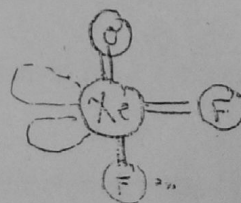
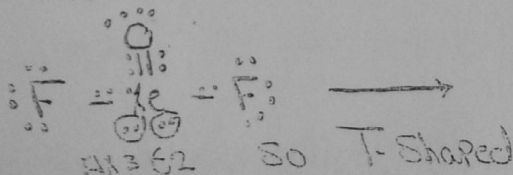
Indicate distortion from regular? ANSWER:

①  $\text{XeOF}_2$

+ Xe is central.

$$\bar{e} = \text{Xe}(18) + \text{O}(6) + \text{F}(14) = 28\bar{e}$$

+ Lewis structure will be

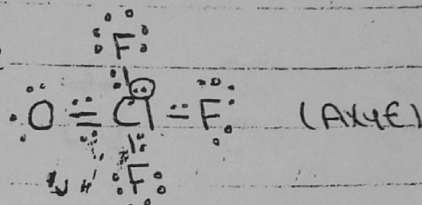


Distortion: Regular Structure is  $AX_5$  (trigonal bipyramidal)  
 But two lone pair of  $e^-$  in equatorial position  $AX_3E_2$  lead to T-Shaped

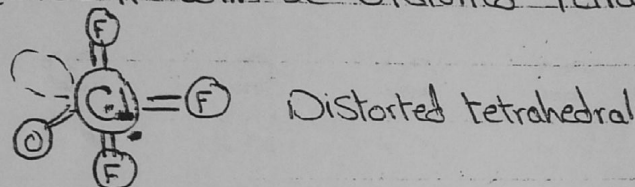
2)  $ClO_2F_3$

$$\bar{e} = Cl(7) + O(6) + F(2 \times 1) = 34e^-$$

Lewis Structure will be



\* Tube  $AX_4E$  So it will be Distorted tetrahedral

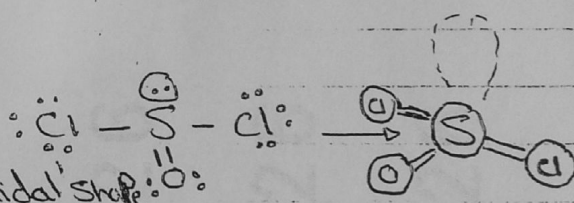


Regular structure is  $AX_5$  (trigonal bipyramidal) but there is one lone pair of  $e^-$  in equatorial position ( $AX_4E$ ) lead to Distorted tetrahedral.

3)  $SOCl_2$

$$\bar{e} = S(6) + O(6) + Cl(4) = 26e^-$$

→ Lewis Structure will be

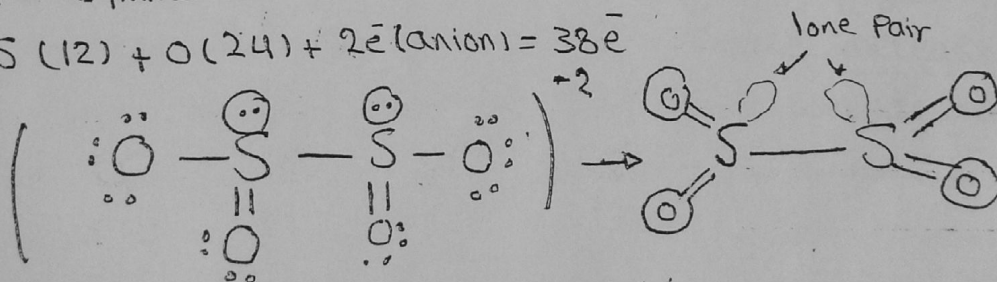


→ Tube  $AX_3E$  So it has a pyramidal shape

→ Regular structure is ( $AX_4$ ) but there is alone pair on (S)  
 So  $AX_3E$  lead to Pyramidal shape not tetrahedral.

4)  $(S_2O_4)^{2-}$  Symmetric S-S

$$\bar{e} = 2S(12) + O(24) + 2e^-(\text{anion}) = 38e^-$$



$O - S = O \rightarrow AX_3E$  So non linear distortion From  ~~$AX_3$~~   $AX_4$



### Question (21)

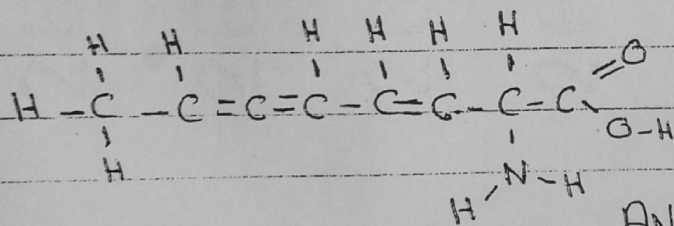
From Li to F the shielding amount of Screening is constant  $\sigma$  (X)  
 From Li  $\rightarrow$  F across 2nd Period  $Z^*$  increases due to increasing atomic number and slight increase in  $\sigma$  so  $Z_{eff} = Z - \sigma$  value so  $Z^*$  become larger, across a period

### Question (22)

22) \* Using the molecular orbital theory, draw molecular orbital diagram for  $N_2$ ,  $NO$ ,  $O_2$ ,  $N_2^+$ ,  $O_2^-$ ,  $F_2$ ,  $O_2^+$  and calculate bond order?

### Question (23)

23) \* By Applying VSEPR theory determine bond angle and type of hybridization around each atom?



ANSWER

C  $\rightarrow$   $sp^3$   $109^\circ 28'$

C  $\rightarrow$   $sp^2$   $120^\circ$

N  $\rightarrow$   $sp^3$   $107^\circ$

### Question (24)

24) Complete:

(Jan 2010)

\* when  $N_2$  goes to  $N_2^+$  the N-N bond distance increase and  
 when  $O_2$  goes to  $O_2^+$  O-O bond distance will decrease